ENTRAL LIBRARY BIRLA CENTRAL LIBRARY PILANI [RAJASTHAN] Chasa No. 547 Book No. W 31P Accession No. 17083

Twentieth-Century Chemistry Edited by J. C. Philip, F.R.S.

PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

TWENTIETH-CENTURY CHEMISTRY

Edited by J. C. PHILIP, F.R.S.

The Theory of Quantitative Analysis and its Practical Application By Henry Bassett, D.Sc., Ph.D.

The Parachor and Valency

By Samuel Sugden, D.Sc., A.R.C.Sc., A.I.C.

A Laboratory Manual of Electrochemistry

By Dr. Erich Müller (Dresden). Translated by H. J. T. Ellingham, Ph.D., A.R.C.S., A.M.I.Chem.E.

The Sorption of Gases and Vapours by Solids

By JAMES WILLIAM McBAIN, F.R.S.

Tautomerism

By John William Baker, D.Sc., Ph.D., A.R.C.S., F.I.C.

Physical Aspects of Organic Chemistry

By WILLIAM A. WATERS, M.A, Ph.D. (Cantab.).

Other volumes in preparation.

PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

 $\mathbf{B}\mathbf{Y}$

WILLIAM A. WATERS

M.A., Ph.D. (CANTAB.)

GORDON WIGAN PRIZEMAN OF CAMBRIDGE UNIVERSITY; LECTURER IN CHEMISTRY IN THE UNIVERSITY OF DURHAM (DURHAM DIVISION)

WITH AN
INTRODUCTION
BY

THE LATE PROFESSOR T. MARTIN LOWRY

C.B.E., D.Sc., F.R.S.

PROFESSOR OF PHYSICAL CHEMISTRY IN THE UNIVERSITY OF CAMBRIDGE

THIRD EDITION



LONDON

GEORGE ROUTLEDGE & SONS, LTD. BROADWAY HOUSE, 68-74 CARTER LANE, E.C.

1942

First Edition - 1935 Second Edition - 1937 Third Edition - 1942

CONTENTS

CHAP.	Preface							PAGE VII
		•	٠		•			V11
	Introduction (by Pre F.R.S.)	OFESS	OR T	. MAF	RTIN	Lowr	Υ,	ix
I.	CHEMICAL AFFINITY	•						I
II.	PHYSICAL THEORIES OF	Moi	.ECUL	AR ST	RUCTU	JRE		13
III.	VALENCY							25
IV.	ELECTRICAL DIPOLES							50
v.	CHEMICAL REACTIVITY							81
VI.	Unsaturation .							99
VII.	FREE RADICALS AND T	HEIR	Non-	ONIC	REAC	TIONS		115
VIII.	IONIZATION AND IONIC	REA	ÇTIONS	5		•		159
IX.	ACIDITY							195
X.	THE REACTIVITY OF H	ALOG	en Co	MPOU	NDS			221
XI.	GENERAL POLARITY					•		238
XII.	HYDROLYSIS AND ESTE	RIFIC	ATION		•	•		262
XIII.	IONOTROPIC CHANGE							288
XIV.	Molecular Rearrang	EMEN	т					324
XV.	Conjugation .	•						372
xvi.	AROMATIC COMPOUNDS-	I						405
XVII.	AROMATIC COMPOUNDS-	–II						453
Index	of Authors							485
SUBJEC	T INDEX							493

PREFACE

A volume reviewing "Physical Aspects of Organic Chemistry" was planned many years ago by Professor T. M. Lowry, F.R.S., who, in 1930, invited the writer to collaborate in the preparation of the projected book, for which much preparatory work had already been carried out. During the past five years the major part of the present volume has been prepared, and discussed frequently, by Professor Lowry and by the writer, who had agreed to act as joint-authors. Late in 1934, however, Professor Lowry regretfully decided that he could no longer take an active part in completing the book, and the writer has endeavoured to fulfil the task alone without deviating from the scheme as originally projected.

Professor Lowry in his "Introduction" to this volume explains the scope of the "physical aspects" of organic chemistry that receive discussion, and, as it will be seen, the intention has been to survey a wide range of chemical theories rather than to devote particular attention to a few specialized themes. The historical aspect of a rapidly-developing subject has been kept continually in view, with the intention of giving a general outline of theoretical organic chemistry rather than one ad hoc point of view.

In concluding, the writer wishes to express his deep indebtedness to Professor Lowry for allowing him the privileges of entering into close collaboration with him, and, secondly, of being granted full permission to make use of the manuscript chapters and notes that he had already prepared.

In addition, the writer begs to thank the chemists of authority to whom Professor Lowry has submitted for

criticism and comment several draft chapters of this book, including, in particular, Professor A. Lapworth, F.R.S., from whom the writer has received, directly, many valuable suggestions.

W. A. WATERS.

THE UNIVERSITY, DURHAM, 1935.

PREFACE TO THE SECOND EDITION

The year which has passed since the first issue of this volume has seen the rapid development of several of the topics which have been discussed in it. Consequently, additional footnotes have been inserted to indicate the trend of the recent research work, and, when necessary, the main body of the text has been corrected.

W. A. W.

Durham,

November 1936.

INTRODUCTION

By the late Professor T. Martin Lowry

One of the most conspicuous features of post-war chemistry has been the application to chemical problems of the new physical conceptions of atomic and molecular structure which were just beginning to emerge before and during the war-period. The enunciation by Helmholtz in 1891 (on the basis of Faraday's Laws of Electrolysis) of the doctrine of the "atomicity of electricity", had been followed by J. J. Thomson's proof that the electron was a universal constituent of all matter, with a constant ratio, e/m, of charge to mass from whatever source it was derived -in marked contrast to the variability of this ratio in the positive ions which were left behind when electrons were stripped from neutral atoms or molecules. The general proposition that matter was electrical in character was, however, not immediately fertile in its application to chemistry—a new Dalton was needed to build up a working theory of the atom of a sufficiently concrete character to explain the familiar facts of chemistry and then to predict others as vet unknown.

This working theory was provided by Moseley's experimental determination of atomic numbers, whereby the charge on the atomic nucleus of Rutherford's atom could be determined, and the number of planetary electrons counted. The classification of these enumerated electrons into groups, sub-groups and grouplets, mainly by means of observations of emission spectra, supplied the necessary detail for the construction of atomic models from which many of the known chemical properties of the elements could be deduced, and a clear interpretation of the periodic classification given.

The physical theory of molecular structure followed rather slowly behind the new theories of atomic structure. The German physicist, Kossel, was one of the first to exploit the relatively-obvious conception of the formation of positive and negative ions by loss or gain of electrons as a means of approach to the stable electronic configuration of an inert gas. This process provided a key to the ionic reactions of inorganic chemistry, and a modern basis for Berzelius's dualistic or electrochemical theory of chemical affinity; but it broke down just as hopelessly as that theory had done in face of the facts of organic chemistry, since it provided no explanation of the formation and stability of symmetrical molecules, such as Cl, or C, H, and could only interpret the formation of a pair of compounds such as CH, and CCl, by the grotesque hypothesis of making carbon a quadrivalent anion in one compound and kation in the other, leading to a perfect reductio ad absurdum in the case of methylene chloride, CH₂Cl₂, where the central atom of carbon would be electrically-neutral and therefore unable to provide any force with which to bind the hydrogen and chloride ions

Not unnaturally, perhaps, the new Dumas was a chemist in the person of G. N. Lewis, who provided a modern basis for the ancient doctrine of substitution in the hypothesis of bond-formation by the process of electron-sharing. The formation of single, double and triple bonds by sharing 2, 4 or 6 electrons, in the form of 1, 2, or 3 DUPLETS, was an integral part of the theory; and the notion that the eight shared electrons of CH₄ or CCl₄, or the eight unshared electrons of neon, formed a cubic octet, was only a temporary aberration, arising from a mistaken geometrical interpretation of an essentially algebraic theory. In view of its obvious incompatibility with van't Hoff's tetrahedral model of the

carbon atom, and the vast range of chemical facts which are now stored up in that model, it could not survive and was soon generally abandoned.

For some time it appeared that the electronic theory of valency could contribute nothing to the fundamental notions of organic chemistry, except an alternative nomenclature, whereby the bonds of Kekulé's formulæ were replaced by pairs of dots, representing duplets of shared electrons. This new nomenclature, however, enabled every bond to be divided into two equal parts, and was therefore used by Lapworth and Robinson as a justification for the notion of partial valencies and in particular of Thiele's theory of conjugation, which could be thrown into a very precise form by using a system of 3-electron bonds, as in J. Thomson's formula for benzene. It also enabled Lapworth in the same year to give an electronic interpretation of his theory of Induced Alternate Polarities.

More direct benefit resulted from the clearing up of the century-old controversy between inorganic and organic chemists as to the polar or non-polar character of chemical affinity. This controversy was finally closed when it appeared that the striving for completeness in electronic systems could manifest itself in two ways, as "electrovalence" by electron-transfer in inorganic or ionic chemistry, or as "covalence" by "electron-sharing" in organic or structural chemistry. This sharp distinction between two types of valency cleared up at once the outstanding problem of the stereochemistry of quinquevalent nitrogen, by establishing the correctness of Werner's analogy of the ammonium ion with the molecule of methane. It also gave a more precise significance to the much earlier suggestion of Lapworth, according to which organic reactions (like inorganic reactions) were believed to proceed through the formation of ions, since this process of ionization could now

¹ Compare W. Hückel, "Theoretische Grundlagen der Organischen Chemie" (1931), vol. 1; p. 25.

² J. Chem. Soc., 1922, 121, 423. ⁸ Ibid., 427.

Loc. cit.

be interpreted as the conversion of a covalence into an electrovalence. The theory of the mixed or semi-polar double bond was also a direct sequel to the doctrine that a single bond may be of either type, since it was then obvious that in a double bond these two forms of single bond might be combined.1 The product was a novel form of linkage, which had not been postulated hitherto,2 but which was soon shown to have a real existence by Sugden's work on the parachor. In the hands of Phillips and his colleagues it also vindicated its utility by providing an explanation of a new type of optical activity, which had not been foreseen and could not have been predicted from the older theories of valency. In this way it was established once for all that the electronic theory of valency has a real contribution to make to the problems of organic chemistry, in addition to providing a new nomenclature and a wonderful clarification of existing ideas.

¹ Lowry, Chem. Age, Oct., 1922; Trans. Faraday Soc., 1923, 18, 285; J. Chem. Soc., 1923, 123, 822.

² The formation of internal salts by compounds such as the betaines and sulphanilic acids had long been admitted by chemists, although the usual confusion between the two forms of valency had found expression in bonds of fantastic extension, which were only displaced slowly as it became clear that the essential properties of these "internal salts" could be expressed satisfactorily only by formulæ such as

The new type of linkage depended on the presence of local charges, comparable with those of the betaines, but attached to contiguous atoms, in compounds such as

This form of bond, with its opposite charges unneutralized, is, however, only stable when each atom carries a complete octet of electrons and is therefore unable (or reluctant) to add to its electronic shell by any further process of electron-sharing. In other cases the semipolar bond can only be formulated as an "activated" form of a non-polar double bond.

The doctrine of the nucleus atom, with its shells of planetary electrons, has, however, a much wider significance in organic chemistry than its application to the mixed double bonds of pervalent compounds of nitrogen, phosphorus or sulphur, since the perfect balance of the positive and negative charges of nuclei and electrons, which is required to produce a completely "non-polar" system, is limited even in organic compounds to a few highly symmetrical molecules. Thus J. J. Thomson, in 1914, showed that the molecules of simple gases and vapours could be segregated into two groups, according as they did or did not carry dipoles of molecular dimensions, the presence or absence of which could be detected (as Debye had shown in 1912) by a study of the dielectric constants of gases. These dipoles necessarily appear whenever the "mass centres" of the positive and negative charges do not coincide, and in general any molecule which does not possess a centre of symmetry (or some other element of symmetry which will bring the two mass-centres together) may be expected to possess a dipole moment of greater or smaller magnitude. Direct measurements have indeed shown that dipole moments exist, not only in compounds, like the betaines or those containing semi-polar double bonds, where we can legitimately picture the molecule as a pair of ions linked together by bonds (the so-called "zwitterion" or multipolar molecule), but also in compounds such as water-vapour or ether, where no such internal ionization need be postulated.

If the formation of ions, either free or bound together by bonds, is accepted as a fundamental mechanism of organic reactions, it is not surprising that these often insignificant dipoles should play a predominating part in determining the course of organic reactions. This is in fact the case, as may be illustrated by Markownikow's rule, since the way in which a halogen hydride is added to a substituted olefine obviously depends on the electronic structure of the substituents, and in particular on the distribution of nuclei and electrons, as manifested by the different dipole moments of these radicals. In a similar way, the peculiar properties

of conjugated systems appear to depend on the possibility of ionizing the successive double bonds, and then recombining the ionic charges between adjacent atoms, in such a way that an electronic charge can migrate from end to end of the system; the direction of migration will then be determined by the pressure of electrons at the ends of the system, so that the dipole-moments of the substituents may again become a predominant factor in determining the course of chemical change.

On these lines a novel type of organic chemistry is being developed in which the course of a reaction is traced back to the electrical properties of the molecule and ultimately to the electronic structure of the component atoms. This development has been rapid and the ideas involved have given rise to no little controversy. Since, however, it is impossible to go back upon, or even to ignore, the knowledge of atomic and molecular structure which has come to us from the physical side during the last two decades, an attempt has been made in the following pages to report upon some of the newer physical aspects of organic chemistry, and to set out some of the questions at issue, even at the risk of inviting criticism, in the hope that more correct ideas than those which we now possess may perhaps emerge from the present period of transition and discussion.

A review of the current position is particularly opportune at the present time, since it has now become clear that, just as there are two principal types of valency, so there are two principal types of organic reactions. Thus, whilst a vast number of reactions are brought about by ionic reagents,' or promoted by ionic catalysts, and may therefore be attributed to the ionization of covalent bonds, the existence of a second group of reagents with an odd number of valency electrons has recently been recognized and made the basis of an alternative system of organic reactions. The "free radicals", which play so important

¹ W. A. Noyes, J.A.C.S., 1933, 55, 656; compare also R. Kuhn, Report of the Fourth Solvay Conference, Brussels, 1930, p. 364. Discussion on "Free Radicals", Trans. Faraday Soc., 1934, 30, 1-248.

a part in photochemical theory and in the discussion of gaseous reactions, have thus found an appropriate place as intermediate products in the familiar processes of organic chemistry. An impartial classification of these processes has thus become possible, in which ions and free radicals can compete on equal terms as active agents in bringing about chemical change.

In writing this volume Dr. Waters has carried to completion a task which the writer of this introduction was unable to fulfil. His exposition is, however, based upon an intimate collaboration over a period of several years, and the views which he expresses are in harmony with those which have been taught and used in the writer's laboratory since the time when Dr. Waters was himself a student at Cambridge. In particular the earlier and more physical chapters were prepared in association with Dr. C. P. Snow, and some of the later chapters were read and criticized by Professor Lapworth, before being passed for publication.

PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

CHAPTER I

CHEMICAL AFFINITY

Berzelius's Dualistic Theory of Chemical Affinity.

During a period of nearly a hundred years, conflicting views as to the nature of chemical affinity have been held by workers in the fields of inorganic and organic chemistry. In view of its earlier development, the first theories of chemical affinity were devised in order to explain the facts of inorganic chemistry, and in particular to explain the opposite properties of acids and bases, and the elimination of these properties by the neutralization of an acid and base to form a salt. These early theories were therefore necessarily dualistic in character. Thus both Lavoisier and Berzelius assigned a unique position to oxygen, and then classified all the other elements according to the character of the products which were formed by their interaction with oxygen. In this way Lavoisier in 1789 classified the non-metals as acid-forming elements and the metals as base-forming elements. The discovery by Nicholson and Carlisle in 1800 of the decomposition of water and by Berzelius and Hissinger in 1803 of the decomposition of salts by the electric current, enabled Berzelius in 1819 to give an electrochemical interpretation to the dualistic theory of By that time the electric current had been Lavoisier. used to isolate the elements sodium, potassium, barium, strontium, calcium and magnesium, and had thus been proved to be the most powerful agent for decomposing

I B

chemical compounds into their constituent elements. The time was therefore ripe for enunciating the doctrine, which came to full maturity a hundred years later in the electronic theories of Kossel and G. N. Lewis, "that in all chemical combination there is a neutralization of opposite electric charges". The dualistic or electro-chemical theory which Berzelius advanced in 1819 was admirably adapted to interpret the structure and properties of acids, bases and salts, which, under suitable conditions, are all capable of being resolved into simpler constituents by electrolysis; but all students of the history of chemistry are familiar with the way in which this theory broke down when applied to the facts of organic chemistry as disclosed between 1815 and 1834, and how the doctrine of substitution or metalepsy which Dumas proclaimed to the Académie des Sciences in January 1834 came to occupy in organic chemistry the same predominant position that Lavoisier's oxide-theory had occupied in the chemistry of salts.

Dumas' Unitary Theory of Chemical Affinity.

Dumas' theory of substitution and the theories of nuclei, types and residues, which quickly developed from it, were essentially unitary in character—a feature that was set out with special clarity by Laurent in 1840 when he described the diatomic molecules of hydrogen and of chlorine, as homogeneous compounds which could undergo double decomposition or substitution to form the heterogeneous compound, hydrogen chloride

$$(HH) + (ClCl) = (HCl) + (ClH).$$

Throughout this period of development, the facts of inorganic chemistry were relegated to the background and the vital contrast between electrolytes and non-electrolytes was either forgotten or ignored. Thus Williamson in 1852 believed that "throughout inorganic chemistry, and for the best-known organic compounds, one single type will be found sufficient; it is that of water, represented as con-

taining 2 atoms of hydrogen to 1 of oxygen, thus $_{\rm H}^{\rm HO}$ ".

In the same way Gerhardt in 1856 classified KOH with C₂H₅OH, and KCl with C₂H₅Cl as examples of the same formula-type; and indeed, throughout the period when empirical formulæ were being developed into structural formulæ, it is impossible to find any admission that alkalis differ in any essential way from alcohols, or salts from esters.

The subsidiary part played by inorganic chemistry, when theories of molecular structure were under consideration, was emphasized during the strenuous years, from 1861 onwards, when Kekulé and others were developing with amazing success the study of the structure of organic compounds; and, when these structures assumed a three-dimensional character, with the enunciation in 1874 by le Bel and van't Hoff of the doctrine of "chemistry in space", the inorganic chemist no longer had any chance of making an effective protest on behalf of dualism in view of the triumphant progress of structural chemistry, based upon a unitary theory of chemical affinity, which characterized the last quarter of the nineteenth century.

The Revival of Dualism.

A strenuous protest against the unitary theory of chemical affinity was made by Werner in 1905. It is noteworthy that Werner's rebellion against the monopoly which organic chemists had held for seventy years in the field of molecular structure, was based upon his introduction into inorganic chemistry at the end of 1802 of the conceptions of "chemistry in space" which van't Hoff had applied to carboncompounds in 1874. There is indeed a remarkably close resemblance between the arguments by which van't Hoff established the tetrahedral configuration of the quadrivalent carbon atom and those by which Werner established the octahedral structure of the radical RA, in the hexammines and related inorganic compounds, and the square configuration of the radical RA, when derived from the octahedral radical RA, by suppressing two of the six groups which were attached in the first instance to the central atom R.

4 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

In Werner's scheme, however, the radicals RA₆ and RA₆ were not necessarily integral molecules, but were usually only the positive (or sometimes the negative) radical of a salt. Thus the octahedral configuration was first suggested in the case of the hexammines

[Ni·6NH₃]Cl₃ [Co·6NH₃]Cl₃ [Pt·6NH₃]Cl₄ in which the radical was present as a bivalent, tervalent or quadrivalent kation, whilst in compounds such as

 $K_4Co(NO_2)_6$ $K_4Fe(CN)_6$ $K_3Fe(CN)_6$

the radical RA₆ was present as a quadrivalent or tervalent anion.

In these compounds it was clear that "the property of the metallic salts which enables them to bind six molecules of water or of ammonia has its ultimate origin in the metallic atom", and that this atom "must possess the power to stand in direct combination with six such groups". These directly-bound groups often included the negative radicals of a simple salt, but these radicals did not then behave as ions, and acquired this character only when displaced from direct combination by the intrusion of molecules of water or ammonia or the like, which possessed a superior power of combination with the central metallic ion. Werner therefore concluded that

"The first condition for the electrolytic dissociation of a salt is the ability of its metallic atoms to combine with a given number of water molecules to form a radical, in which we must think of the water molecules as arranged in such a manner that a direct combination between the metallic atom and the acid residue can no longer take place."

Ionogenic and Non-Ionogenic Linkages.

These novel ideas, which Werner set out in fuller detail in his Neuere Anschauungen (1905), make it clear that valency can manifest itself in two ways. There is, on the one hand, a 'primary valency', which results in the formation of salts such as NiCl₂, CoCl₃, PtCl₄. The halogen is here

¹ Werner, Zeit. anorg. Chem., 1893, 3, 287, 282, 296.

held by an ionogenic linking, which enables it to split off as a negatively-charged anion from the metal, which is thus converted into a positively-charged kation; but this process of ionization (already rendered familiar by the theory of electrolytic dissociation which Arrhenius had introduced in 1887) was greatly facilitated by the action of the "secondary valency", by which the metal was able to attach itself to integral molecules of water or ammonia by means of non-ionogenic linkages. In this way Werner was led to discriminate, although not always with perfect clarity, between two different types of valency. He also drew a sharp distinction between the valency-number and the co-ordination-number of an atom as follows:

"The valency-number gives the maximum number of univalent atoms which can be united directly to the atom without the intervention of other elementary atoms."

"The number of atom-groups with which an elementary atom co-ordinates to form a compound radical, in the manner described above, I might call the co-ordination-number of the atom in question."

"The co-ordination-number gives the maximum number of atoms which can stand in direct union with the atom in question." 1

The difference between the valency-number and the co-ordination number was illustrated as follows: In carbon, the values of these two numbers are equal; in almost all other elements, however, they are different from one another. Thus boron and nitrogen are both tervalent, as in the hydrides BH₃,² NH₃, and in the chlorides BCl₃, NCl₃. Since, however, the co-ordination-number of both elements is four, they can form complexes of the type MX₄, but only with the help of a third element. Thus, the addition of a fourth atom of fluorine to BF₃ involves the addition of a

¹ Werner, Zeit. anorg. Chem. 1893, 3, pp. 326, 328.

² Later experiments have shown that boron is *not* tervalent in its hydride, since this has the molecular formula of *boroethane*, B_2H_6 (compare C_2H_6), and evidently contains quadrivalent boron, although there is an unexplained shortage of two electrons in the molecule as compared with ethane.

6 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

fluoride ion, which must be accompanied by a metallic kation, as in the equation:

$$BF_3 + KF \longrightarrow [BF_4]^- K^+$$
;

Conversely, the addition of a fourth atom of hydrogen to ammonia, or of a fourth positive radical to a tertiary amine, can only be effected with the help of an anion, as in the equations:

$$NH_3 + HCI \rightarrow [NH_4]^+ CI^-$$

 $NR_3 + RI \rightarrow [NR_4]^+ I^-$

Werner's Theory of Valency.

As a result of his work on the molecular structure of inorganic compounds, Werner came to the conclusion that the theory of chemical affinity, as it had been developed by Kekulé and van't Hoff, was fundamentally unsound. These prevailing views he summarized as follows:

"The fact that an atom is able to combine with a definite maximum number of other atoms is even to-day, in many cases, interpreted as meaning that the affinity of an atom only comes into action through a definite number of units of valency which are considered as independent forces. According to this conception the unit of valency consists of a fraction of the affinity of the atom appearing as an independent force, and acting so as to hold together two atoms, thus forming a molecule. Some experimenters (Erlenmeyer, Lehrbuch der organischen Chemie, p. 40) further assume that these independent forces (units of valency) act only from definite positions on the surface of an atom, and that multivalent atoms therefore have their units of valency separated and assigned to definite points on their surfaces. Other workers (van't Hoff, La Chimie dans l'espace, 1875) have added the hypothesis that the units of valency can act only in definite directions in space. Thus, for example, the four valencies of carbon are supposed to act in the direction of the four corners of a regular tetrahedron, the carbon atom being at its middle point."1

In opposition to van't Hoff's conception of individual directed valencies, Werner appears to have adopted the

views of Lossen, who states that "The valency of an atom is [merely] a number which expresses how many atoms are present in its combining zone" and of Claus, who says that "The hypothesis that valency in multivalent atoms is a pre-existing force acting in definite units of affinity is as unfounded as it is unnatural." Werner's own views are expressed in the following sentences.

"Let us suppose that the atom consists of homogeneous matter. Hypotheses on the form of the atom are unnecessary . . . but for the sake of simplicity we can suppose it to be a sphere."

"Affinity is an attractive force which acts from the centre of the atom, and is of equal value at all points on its surface. With such a definition of affinity it naturally follows that

separate units of affinity do not exist."

"Valency is a number, which has been empirically found . . . and which expresses the ratio in which atoms combine with one another. It does not depend upon one atom alone, but rather upon the nature of all the atoms present in a particular molecule."

"The amount of affinity saturated by the linking up of two atoms is distributed on a definite circular segment on the surface of the atom, and varies within wide limits with the

nature of the atom."

Reconciliation of the Unitary and Dualistic Theories of Valency.

The struggle between the unitary and dualistic theories of chemical affinity, which raged for nearly a century, was a form of civil war between inorganic and organic chemists. It was finally resolved in modern times with the help of those workers, mainly physicists, to whom we are indebted for our present knowledge of atomic and molecular structure. Their work made it obvious for the first time that, although chemical affinity in all its manifestations has a single origin in the somewhat mysterious forces which form the subject of "quantum theory", there are two ways in which it may exert itself. One of these, depending

¹ Ann. Chem. Pharm., 1880, 204, 327.

² Ber., 1881, 14, 432. ³ "New Ideas", p. 73.

on "electron-transfer" (p. 10), leads to the union of elements of opposite type, as in the dualistic theory of the inorganic chemists; the other, depending on "electronsharing" (p. II), is concerned with the union of like elements, which plays so important a part in the construction of the carbon chains and rings of organic chemistry, and which found its expression in the doctrine of substitution and the unitary theory of chemical affinity. Both types of chemical affinity have, therefore, a real existence, and can be clearly interpreted—if not predicted—by means of the ELECTRONIC THEORY OF VALENCY (p. 10), but each is only applicable over a restricted range. It was, therefore, iust as wrong for organic chemists to apply the "directed valency" theory of van't Hoff to mineral salts, as it was for Berzelius to try and force the benzovl radical into the mould which had served so well in the case of rock salt or potash alum.

The Electronic Structure of Atoms and Molecules.

It is unnecessary at this stage to set out in detail the views which are now held as to the nature of atomic and molecular structure. It will suffice, therefore, merely to refer to Rutherford's nucleus atom, as a system in which a compact central nucleus carrying a positive charge $+ \mathbf{Z}$, is surrounded by an equivalent number Z of negativelycharged electrons. The nucleus is itself believed to be a composite structure consisting of *protons*, or hydrogen nuclei, with a mass of 1.67×10^{-24} grams and a positive charge of 4.774×10^{-10} e.s.u., and of electrons with a mass 1,847 times smaller, but with an exactly equivalent negative charge. The protons are always in excess of the nuclear electrons, and this excess determines the atomic number of the atom, which may be defined either by the magnitude of the resultant positive charge on the nucleus, or by the number of planetary electrons, which must be associated with the positively-charged nucleus in order to convert it into a neutral atom.

Quantum-theory must, then, be invoked to account for

the extraordinary stability of the planetary electrons ¹ in those atoms of which the atomic number is 2, 10, 18, 36, 54 or 86, as contrasted with all other atoms. These maxima of stability form a simple algebraic series, generally known as the *Rydberg series*, which may be set out with the symbols of the corresponding elements as follows:

A detailed study of emission and absorption spectra (including X-ray as well as visual spectra) has resulted in the separation of the *groups* of electrons which make up the Rydberg series into *sub-groups*, so that the 54 electrons of xenon can be split up as follows:

$$2 + 8 + 18 + 32 + 8$$

= $2 + (2 + 6) + (2 + 6 + 10) + (2 + 6 + 10 + 14) + (2 + 6)$

By the same methods we can also carry the sub-division down to still smaller grouplets, each containing a small even number of electrons. We can then account for the limitation of these numbers by means of Pauli's "exclusion principle" (see below), and thereby secure a glimpse of the fundamental factors on which the periodicity of the elements, as discovered by Mendeleef and Lothar Meyer, is based; but, for the purpose of the present exposition, the numbers of the Rydberg series are sufficient as defining the con-

¹ The relative stability of the *nuclei* is a separate problem, the solution of which depends on a study of their spontaneous disintegration in radioactive elements and of their artificial disintegration by bombardment with nuclei of some of the lighter elements.

² Pauli's **exclusion principle** states that no two electrons in a given atom can have the same four **quantum numbers**, n, l, j, m, where

the **principal quantum number** n is characteristic of the group of 2, 8, 18, 32 electrons.

the **subsidiary quantum number** *l* is characteristic of the subgroup of 2, 6, 10, 14, etc., electrons.

the inner quantum number j is characteristic of the grouplet of 2, 4, 6, etc., electrons.

the magnetic quantum number m is characteristic of the individual electrons in a grouplet.

ditions under which a maximum of stability may be expected in the atoms of an element of given atomic number.

The Electronic Theory of Valency: Electrovalence and Covalence.

In the case of compounds, similar maxima of stability are to be expected when each atom, by hook or by crook, has acquired the number of planetary electrons which are characteristic of an inert gas, as represented by one member of the Rydberg series of numbers. The two ways in which stable groupings of electrons may be acquired were both suggested by Sir J. J. Thomson in 1904 1 at a time when the atomic numbers of the elements were still unknown and the Rydberg series had therefore not yet been formulated; they are (a) electron-transfer and (b) electron-sharing.

(a) Electron-transfer. The alkali metals, which follow the inert gases in the periodic classification of the elements, contain one electron more than the Rydberg number; the halogens, which precede the inert gases, contain one electron less. When, therefore, an atom of potassium ($\mathbf{Z} = \mathbf{19}$) and an atom of chlorine ($\mathbf{Z} = \mathbf{17}$) are brought together, both atoms can acquire the electronic grouping of argon ($\mathbf{Z} = \mathbf{18}$) by the simple process of transferring an electron from potassium to chlorine and thus using the surplus electron of the metal to make up the deficit of the nonmetal. In this process the metal acquires a positive charge and becomes a kation, whilst the non-metal acquires a negative charge and becomes an anion.

Nuclear charge . Planetary electrons .	•	K 19 19	+	Cl 17 17	=	† K 19 18	+	Cl 17 18
					-			
Difference	•	0		0		+ 1		<u> </u>

The completed groups and sub-groups of electrons which are thus formed are of very great stability. The ions thus produced are therefore more stable than the atoms from which they were formed, and do not give up their electrical charges even when brought into intimate contact with one another. Their opposite electric charges, however, draw them together so strongly that the aggregates of potassium and chloride ions, which make up the crystals of solid potassium chloride, do not melt until the temperature is raised to 775° C.

This type of linkage was discussed fully by Kossel in 1916, and was described by Langmuir in 1921 as electrovalence. It has all the characteristics of the type of chemical affinity postulated by Werner, since the electrostatic attraction exerted by each ion is distributed equally round the ion, when it is remote from all other ions, and can be subdivided into arbitrary fractions according to the number of oppositely-charged ions by which it is surrounded. Thus univalent chlorine is surrounded by 4 atoms of univalent silver in silver chloride, by 6 atoms of univalent sodium in rock salt, and by 8 atoms of univalent cæsium in cæsium chloride. The indivisible single bond postulated by the chemists of the early fifties has, therefore, no real existence in these compounds; and even the vapour of sodium chloride may be regarded as consisting merely of ion-pairs, in which the only form of "bond" between the two atoms is an electrostatic field radiating from one atom and converging more or less completely upon the other.

(b) Electron-sharing. One of the most striking characteristics of non-metallic atoms is their ability to combine, not only with metals, according to the scheme of Berzelius, but also with a second atom of identical type, in utter defiance of that scheme, to form diatomic molecules of extraordinary stability. This type of combination can be interpreted as depending on a process of electron-sharing, whereby two atoms are held in common by both nuclei, just as two carbon atoms may be held in common by two aromatic rings, with the result that only 10 atoms of carbon are required to produce a binuclear molecule in naphthalene as contrasted with 12 atoms in diphenyl. By a similar process of sharing, the 34 electrons of two atoms of chlorine can provide a complete shell of 18 electrons for each nucleus.

This type of union was discussed fully in 1916 by G. N. Lewis, and was described by Langmuir in 1921 as covalence.

The two shared electrons were described by Lewis as a duplet, and were identified with the bond of the organic chemist. The idea that two electrons are required to form a bond is in accord with traditional theories of valency which represent the bond as using up one unit of affinity of each of the atoms that unite. In practice it leads to the presence of an even number of electrons in the molecules of nearly all stable chemical compounds (the principal exceptions being found in NO, NO2 and ClO3), since in this way the molecule secures complete electron-groups like those which lead to maxima of stability in the atoms. Later developments of the electronic theory have emphasized the value of pairs of shared electrons by attributing the stability of the two-electron bond to the coupling of two spinning electrons oriented with spins in opposite senses: and the theory of wave-mechanics has added further value to this combination by assigning a definite exchange-energy to the coupling of two electrons which are in every respect indistinguishable from one another.

The process of electron-sharing gives rise to a linkage having all the characteristics postulated by Kekulé and van't Hoff. Thus the number of valencies is limited to the number of vacancies which can be filled by the sharing of electrons; this number is in general identical with the negative electrovalence of the element, as measured by the number of electrons which it takes up when converted into an anion. Moreover, since the electron is indivisible, a covalence cannot be subdivided like an electrovalence; and, finally, since the orbits of the electrons must be oriented in some regular way relatively to one another, we also arrive at the scheme of directed valencies postulated by van't Hoff. In this way, then, it becomes clear that the electronic theory not only indicates the existence of two types of valence, but actually points to the existence of exactly those characteristics which formed the subject of such prolonged controversy between inorganic and organic chemists.

CHAPTER II

PHYSICAL THEORIES OF MOLECULAR STRUCTURE

Physical Theories of Molecular Structure.

The physical theory of the structure of molecules has been developed much less fully than that of simple atoms or ions, and it is not yet possible to speak so dogmatically about the behaviour and properties of electrons which are held in common by two nuclei, as of those which are under the control of a single nucleus. On the other hand, experimental evidence bearing on the physical structure of molecules is being steadily accumulated, and interpretations based upon the theory of quantum mechanics are being developed. As a result, the electronic structures of some simple molecules are now known with considerable accuracy, whilst those of more complex structures may be surmised.

The experimental evidence, upon which our knowledge of the electronic configuration of molecules rests, has been provided almost entirely by the study of band spectra. These are produced by excited molecules just as line spectra are produced by excited atoms, but are much more complex in their structure. Thus, according to Bohr's theory of line spectra, each line depends on the change of energy resulting from an electron jumping from one permitted orbit to another. Since the number of permitted orbits is limited, line spectra are relatively simple, although in some cases the number of lines may be considerable. In band spectra, however, each line is spread out into a band or series of flutings as a result of abrupt changes of vibrational energy accompanying the electronic changes, since a pair of nuclei which are joined by a bond can vibrate to and fro in a way

¹ Phil. Mag., 1913, [vi] 26, 1, 476, 857.

that is obviously impossible in the case of single atoms or ions. Moreover, small changes occur in the rotational energy of the molecule, as a result of which the flutings of the bands are resolved into series of fine lines, in a way which is again impossible in the case of simple atoms, which behave from this point of view as point systems. Band spectra are therefore very complicated, and it was not until 1924 that Victor Henri, following up the earlier suggestions of Heurlinger showed how chemists could interpret them by making use of the fact that the rotational energy and the vibrational energy of a molecule may both be quantized, i.e. restricted to a regular series of permitted increments, just as Bohr's theory limits the electronic changes to a series of stationary states or energy levels of the atom.

(a) The first important advance in the study of band spectra was due to a suggestion made independently by Mecke 3 and Birge 4 and developed extensively by Mulliken.⁵ This was to the effect that an analogy might be found between the spectrum of a molecule and that of an atom with an equal number of outer electrons. A simple example is found in the doublet character of the sodium spectrum, which is so familiar in the case of the D-line. This is characteristic of systems with a single "emitting" or "unbound "electron and is therefore observed in all the alkali metals. It has been found also in a number of unstable aggregates, BeH, MgH, CaH, BO, CN, etc., which cannot be isolated, but whose spectra are more easily observed in a discharge-tube than those of more stable molecules. These aggregates, then, must also contain a single "emitting" electron, like the alkali metals.

This fact enables us to draw one very important deduction, which had been established hitherto only on an empirical basis, namely that the bond between the atoms is formed by a duplet of shared electrons. This deduction may be illustrated

¹ Victor Henri and de Laszlo, Proc. Roy. Soc., 1924, 105, 662.

² Dissertation (Lund), 1918.

⁸ Zeit. für. Physik, 1924, 28, 261. ⁴ Nature, 1926, 117, 300. ⁵ Phys. Rev., 1925, 26, 561.

by reference to the spectrum of MgH. In this case, the electrons of the component atoms of magnesium and of hydrogen are distributed as follows:

Since the "multiplicity" of the lines is one more than the valency of the atom, the magnesium spectrum consists of triplets, whilst that of hydrogen consists of doublets. If, therefore, all the three outer electrons remained free in magnesium hydride, MgH, the spectrum should consist of quadruplets. Actually it consists of doublets, since two of the outer electrons have been used up or "bound" in forming the bond between magnesium and hydrogen, and are no longer available as "emitting electrons". On the other hand, NH, with (2+5)+1 electrons, resembles oxygen, with (2+6) electrons, and gives rise to a similar spectrum.

(b) A theoretical basis for the empirical view that a bond is formed by a duplet of coupled electrons was provided in 1927, by Heitler and London, who were able to describe the formation of a hydrogen molecule from two hydrogen atoms in terms of quantum mechanics. The results of their calculation may be expressed graphically as shown in Fig. I. on page 16.

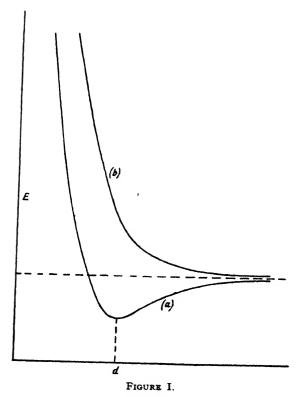
The hydrogen atoms are first imagined to be at infinity. The variation of potential energy E of the system as the interatomic distance d gets less then follows one of two paths, since it may either ascend smoothly as the distance grows less, or it may pass through an intermediate minimum. Thus, in Fig. I, curve (a) represents the formation of a stable molecule, with the atoms separated by a distance corresponding to the position of the minimum. Curve (b), with no position of minimum potential energy, represents a mutual repulsion of the atoms at all distances from one another.

¹ Zeit. für. Physik, 1927, 44, 455.

² Curves of this kind can be deduced by postulating an attraction varying as d^{-7} , and a repulsion varying as d^{-18} . (Lennard-Jones, *Proc. Phys. Soc.*, 1931. **43**, 476.)

16 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

The difference between these two possibilities is determined by the nature of the functions which have to be used in the calculation and can be interpreted in terms of the "spinning electron" of Goudsmit and Uhlenbeck.¹ Thus



curve (a), which indicates the formation of a molecule, corresponds to a condition in which the spins of the electrons of the two hydrogen atoms are oriented in opposite

¹ Physica, 1926, 6, 273. The hypothesis of the spinning electron was introduced because the three degrees of freedom of a point-electron provided only half the number of energy-levels that were required to explain the spectra of the elements, and allowed for only half the number of elements in the periodic classification.

senses, whilst curve (b) corresponds to a condition in which the spins are in the same sense. The covalent bond in a hydrogen molecule must, therefore, consist of two electrons of opposite spin, just as the duplet of K-electrons in an atom consists of two electrons of opposite spin.

This extremely important result is a theoretical justification for the empirical description of the duplet bond as given by Lewis, or as derived from band spectra, since there is little doubt that the same conditions must apply to most of the covalent bonds known to the chemist. Approximate theoretical treatments of covalent bonds in molecules more complicated than hydrogen have indeed already been made by Pauling and Hund, and also by Lennard-Jones, 1 In all this work the covalent bond is assumed to consist of two electrons, opposite in spin, contributed from each of two atoms. The opposition of spin is therefore now regarded as an essential condition for the formation of a bond, although the coupling of the magnetic fields of the spinning electrons provides only a small part of the energy of bond-formation. The greater part of the binding force is provided by the electrostatic attractions of the nuclei and electrons, but these can only operate when the opposition of the spins of the electrons permits the molecule to be formed.

- (c) Lennard-Jones has discussed the results described in (a) and has drawn an important distinction between electrons which are controlled by a single nucleus and are said to be in an atomic orbital, and electrons which are controlled by more than one nucleus and are said to be in a molecular orbital. This distinction has the advantage of emphasizing the contrast between the behaviour and properties of mononuclear and binuclear electrons. Two applications may be given to illustrate the usefulness of this method of approach.
- (i) The four valency-electrons of carbon are of two kinds.² These are distinguished as s and p electrons, according as the subsidiary quantum number has the values l = 0, l = 1. In methane, eight electrons are shared between one atom

¹ See, for instance, Trans. Faraday Soc., 1934, 30, 70.

² See also p. 43.

т8

of carbon and four atoms of hydrogen. Is it necessary to postulate that the four bonds must be of two types, according as they are derived from an s or p electron of carbon; or are they really all alike? In the nomenclature of Lennard-Jones and Mulliken the two K-electrons of the inner shell of carbon may be relegated to "atomic orbitals"; but the other eight electrons are transferred to "molecular orbitals". In this condition they have ceased to be s and p electrons and are therefore at liberty to form four identical duplets, if the little-known rules of the game of molecule-building allow it.

(ii) Pauli's exclusion principle limits the number of electrons in the K-shell of an atom to two; and this limitation is observed also in molecular hydrogen, where two electrons are shared by the two nuclei. These electrons, however, are no longer in an "atomic orbital" but in a "molecular orbital", and it would be very dangerous to invoke the exclusion principle as evidence that the molecular orbitals associated with a hydrogen atom can never hold more than two electrons. There is, therefore, no valid objection on this account to postulating that, in the acid-fluoride ion of KHF₂, the hydrogen nucleus may share four electrons with two atoms of fluorine :F:H:F:, instead of sharing only two electrons with one atom of fluorine as in H:F: 1.

Ionization of Molecules.

The view is very widely held, both by chemists and by physicists, that no sharp distinction can be drawn between a covalence and an electrovalence. On the theoretical side it is impossible to be dogmatic, but as a matter of practical convenience it is assumed throughout the present volume that the two types of valency exist and can be distinguished from each other.² In terms of the theory of electron-orbits

¹ See p. 34 for a fuller discussion of the "hydrogen bond".

² See N. V. Sidgwick, "The Covalent Link in Chemistry" (Cornell Univ. Press, 1933), Chapter 2, for a justification of this point of view.

this means that the orbit of a given electron envelops either one nucleus or two, and that there is a definite quantumjump involved in the transition from a highly-deformed orbit around one nucleus to an orbit which is quantized around two nuclei. In terms of Lennard-Jones's nomenclature it means that the transfer from an "atomic orbital" to a "molecular orbital" takes place by integral electrons and is not a mere continuous leakage of electron-stuff from one condition to the other.

On the practical side it has been urged that a complete series of intermediate types exists, e.g. between rock salt and methane, which makes it impossible to delimit a boundary between the two forms of chemical affinity. It is possible that valid examples of intermediate types may exist, but most of the cases that have been cited for this purpose are singularly unconvincing, since they usually depend either on a mere statistical gradation in the percentage of ionization or on a progressive change in the energy of ioniza-These gradations can both be predicted as corollaries to a theory of clean-cut ionization, and cannot, therefore, be used as evidence against such a theory. This contention. can be illustrated, for instance, by examining the series:

- (i) KCl, (ii) HCl, (iii) HgCl, (iv) HgC, (v) H,O, (vi) CH,
- (i) and (vi) There would be widespread agreement for the view that potassium chloride (i) is completely ionized and that methane (vi) consists of covalent molecules, although a few chemists may still long for a real molecule of potassium chloride, and a few physicists may still think that methane is adequately represented by Kossel's ionic formula.
- (ii) Hydrogen chloride is the best possible example of an acid-forming hydride. As W. L. Bragg has pointed out, however, the hydrogen ion H differs from all other ions in being a naked nucleus—a proton with no electrons either in the nucleus or outside it. It can therefore only exist in a transitory form in a vacuum, since in presence of other ions or molecules there is nothing to prevent it from penetrating the electron-shell of the first atom that it meets until

repelled by the positive charge of the nucleus.¹ It will then remain there, provided that there are two unshared electrons in whose orbits it can wrap itself—since the exclusion principle may perhaps be applied to nuclei as well as to electrons. The molecule of hydrogen chloride has a large dipole moment, but has the same type of crystal-structure and almost the same volume as solid argon. According to the views set out above, it is wholly covalent, and only undergoes ionization when some other molecule is present with unshared electrons to which the proton can transfer itself, e.g.

$$OH_2 + HCl \rightleftharpoons \overset{\dagger}{O}H_3 + \overset{\dagger}{Cl}$$
 or $NH_3 + HCl \rightleftharpoons \overset{\dagger}{N}H_4 + \overset{\dagger}{Cl}$
but not $HCl \rightleftharpoons \overset{\dagger}{H} + \overset{\dagger}{Cl}$.

- (iii) and (iv) In the case of mercuric chloride and mercuric cyanide, we are again dealing with covalent molecules, the formation of which by metallic mercury is established clearly enough by the recognition of "chain molecules" in the crystal-structure of calomel, Cl—Hg—Hg—Cl. The ionization of these salts is therefore reversible, Cl—Hg—Cl \rightleftharpoons Cl + Hg + Cl, instead of being complete as in the alkali halides, and their feeble conductivity depends merely on the smallness of the proportion of molecules which are ionized.
- (v) Water is another example of a molecule which is covalent but ionizable. Ionization proceeds according to the equation $2H_2O \rightleftharpoons H_3O + OH$, but the tendency to ionization is even less pronounced than in mercuric chloride. The ions are, however, entirely distinct from the molecules, since ionization involves the transfer of a parasitic proton from one host to another, and it is impossible to formulate

¹ Thus Fajans (*Ber.*, 1919, **21**, 549 and 709) has calculated that the heat of hydration of a proton to form an aqueous hydrogen ion is 260,000 calories per gram-equivalent. From this number it has been deduced that the theoretical concentration of free protons in water is 10^{-150} !

² See p. 198.

the atom-groups OH, OH₂ and OH₃ as the products of a continuous migration of proton-substance. In none of these compounds, therefore, can the reversibility of the ionization be accepted as evidence of the obliteration of the boundary between molecules and ion-pairs, or of a progressive conversion of a covalence into an electrovalence.

Polarization of Molecules and Ion-Pairs.

The distinction between covalence and electrovalence may be blurred (but not obliterated), not only by the statistical effects of reversible ionization, but also by the influence of polarization. As a limiting case, we may consider a diatomic molecule either of hydrogen or of chlorine. These molecules are represented by symmetrical formulæ, in which the mass-centres of the nuclei and of the electrons coincide, so that the molecule has no dipole moment (see Chapter IV). Even under the most favourable conditions, however, this is merely a statistical average, since temporary dipoles must be created whenever an unsymmetrical oscillation of the positive and negative components takes place. More important still is the distortion resulting from a "polar" environment, such as may be produced by the stray fields from adjacent ions, ion-pairs or ionic aggre-These fields will distort the electron orbits or (in terms of wave-mechanics) alter the distribution of electrondensity around the two nuclei. The molecule will therefore become "polarized" and unsymmetrical, and the additional energy required to ionize it may be reduced almost to the vanishing-point without disrupting the molecule. There can, for instance, be little doubt that the action of potassium hydroxide on chlorine, proceeds by a process of this kind, in which the molecule of chlorine is first polarized and then disrupted by the ions of the alkali:

$$\ddot{K}$$
 + Cl—Cl + $\ddot{O}H$ = \ddot{K} + $[\ddot{Cl}$ + \ddot{Cl}] + $\ddot{O}H$ = $\ddot{K}\ddot{Cl}$ + ClOH.

The final rupture of the bond is, however, a definite step, which is quite distinct from the mere polarization of the molecule.

Polarized molecules may also be produced by linking together unlike atoms, e.g. of hydrogen and chlorine, whereby a permanent separation of the mass-centres of the positive and negative charges ensues. In this way it is possible for a covalent molecule to possess a large dipole moment, although this will be less than if the molecule were disrupted into ions. Conversely, the dipole moment of an ion-pair such as silver iodide, AgI, may be reduced below the value calculated from the radii of the atoms and the charges which they carry, either because the radii may be variable or because the two ions may alter the distribution of each other's charges. These effects are attributed to deformation of the ions, but do not suffice to convert them into covalent molecules. Nor can feeble electrolytic conductivity be cited as evidence of the formation of real molecules, since ionization and ionic dissociation are clearly two distinct operations. Thus, ion-pairs, especially if highly deformable, will not be expected to dissociate unless the solvent possesses the power to penetrate between the ions and to wrap them round so effectively as to prevent them from coming in contact with one another. This condition was foreseen clearly by Werner, as long ago as 1892, but it is only now beginning to be generally admitted. Experimental support may be found in the work of A. R. Martin,² who has shown that, when the measurements of conductivity of electrolytes dissolved in benzonitrile have been corrected for the effects of electrostriction by the formula of Debye and Hückel, the degree of dissociation is much less than 100 %, even in the case of metallic salts which appear to crystallize in ionic aggregates.

Intramolecular Ionic Charges.

It will be convenient at this stage to refer to the question of localized electric charges in covalent molecules. (i) Local charges are obviously produced whenever the covalent

¹ Fajans, Rocz. Chem., 1926, **6, 395**; Z. Krist, 1928, **66,** 321. ² I. Chem. Soc., 1928, 3270.

molecules of an acid are converted into the ions of a metallic salt, e.g.

$$CH_3 \cdot CO \cdot OH \rightarrow CH_3 \cdot CO \cdot O \stackrel{+}{Na}$$
 $CH_3 \cdot CO \cdot OH \rightarrow CH_3 \cdot CO \cdot OH \stackrel{+}{Na} \stackrel{-}{O} \stackrel{-}{Na} \stackrel{-}{$

The essential operation in this act of neutralization is the removal of a proton from a hydroxyl-radical of the acid under the influence of the hydroxyl-ion of an alkali. This leaves behind a surplus of one electron, since a duplet, formerly shared between oxygen and hydrogen, now belongs exclusively to oxygen. This primary change will certainly be followed by a secondary displacement of positive and negative charges; but it is convenient to assign the whole of the surplus charge to the oxygen and to represent the

univalent oxygen atom by the symbol -0. The same result can be reached by noticing that the univalent oxygen atom, with a nuclear charge of 8, carries two K-electrons, six unshared electrons and half of a shared duplet. The total of nine leaves it with a surplus charge of minus one.

(ii) On passing from the silicate ion to the phosphate, sulphate and perchlorate ions, the positive charge on the nucleus is progressively increased from Z = 14 to 17.

In this process the total external field of the ion is reduced by I, 2 or 3 units, so that the phosphate ion is tervalent, the sulphate ion is bivalent and the chlorate ion univalent; but the quadricovalent chlorine atom, which shares four pairs of electrons with oxygen, may still be represented with a triple positive charge, since it has only 2 + 8 inner electrons and a half-share of eight more, making a total of 14, as contrasted with a nuclear charge of 17. This device of assigning integral electric charges to individual atoms, which was suggested by Lowry in 1922,1 has recently

¹ Trans. Faraday Soc., 1923, 18, 285.

24 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

been used by Pauling ¹; it is concerned with the number and not with the form of the electron orbits around each atom, and is therefore not vitiated by any distortion of the orbits or compensation of dipoles which may occur as a sequel to the development of localized negative and positive charges on passing from silicic acid to the perchlorate ion.

¹ J.A.C.S., 1931, **53**, 1367.

CHAPTER III

VALENCY

Covalence and Electrovalence.

In a preceding chapter (Chapter I, pp. 10-12) the two principal types of valency have been discussed. It now remains to describe some of the variants on these principal types that have been suggested from time to time. These include:

- (a) The mixed or semi-polar bond.
- (b) The dative bond or co-ordinate link.
- (c) The hydrogen bond.
- (d) Odd-electron bonds.
- (e) Bivalent carbon.
- (f) The oxygen bond.

Polarity of Double Bonds.

As indicated above (p. 18), the link between two univalent atoms must be either a covalence, as in Cl—Cl, or an electrovalence, as in K Cl. In the case of a double bond, however, *three* options are available, since the two atoms may be linked by

(i) a double covalence as in acetone or ethylene

$$CH_{3}$$
 $C=O$ H $C=C$ H

- (ii) a double electrovalence as in calcium sulphide Ca S.
- (iii) one covalence and one electrovalence as represented by the symbol $\overset{+}{X}-\!\!\!-\!\!\!-\!\!\!\!-\bar{Y}.^1$

Lowry, J. Chem. Soc., 1923, 123, 822. As an alternative, the symbol \Rightarrow was suggested, where the barb shows the direction of the field of force from + to -; but this symbol becomes ambiguous in the case of ions, where the positive and negative charges are no longer equal in number.

The possibility of thus superposing an electrovalence on a covalence was first suggested in a paper on "The Polarity of Double Bonds ",1 which was read in October 1922, but not published until the spring of 1923. The view was then expressed that "a double bond in organic chemistry usually reacts as if it contained one covalency and one electrovalency". Bonds of this character were described as "mixed double bonds". Thus the reactive form of acetaldehyde was written as CH₃·CH—O and that of ethylene as CH₃—CH₃, where each minus sign indicated an excess of one planetary electron above the net nuclear charge and each plus sign a deficit of one electron. In the same way, the reactivity of chlorine and of methyl iodide was attributed to "the readiness with which a covalency gives rise to an electrovalency, as in the balanced equations

$$Cl_2 \rightleftharpoons \stackrel{+}{Cl} + \stackrel{-}{Cl}$$
 and $CH_3I \rightleftharpoons \stackrel{+}{CH}_2 + \stackrel{-}{I}$ ".

Evidence in favour of this theory of polar activation was supplied in July 1923 when Norrish showed 2 that the interaction of ethylene and bromine (and subsequently 3 of ethylene and chlorine), which was already known to be a surface reaction,4 can be arrested by enclosing it in a vessel lined with paraffin wax, but is promoted to an increasing degree by contact with cetyl alcohol, stearic acid and wet glass (p. 172). Polar activation can be effected even more readily by a polar reagent than by a polar catalyst, as, for instance, when acetone undergoes polar activation under the influence of potassium cyanide in the cyanhydrin reaction ⁵ (p. 104).

$$\begin{array}{ccc}
CH_{3} & C=0 & \rightleftharpoons & CH_{3} & \longleftarrow & \overline{O} \\
CH_{3} & \longleftarrow & \overline{O} & + & \longleftarrow & CH_{3} & \longleftarrow & \overline{O} & K \\
CH_{3} & \longleftarrow & \overline{O} & + & \longleftarrow & CH_{3} & \longleftarrow & \overline{O} & K
\end{array}$$

¹ Lowry, J. Chem. Soc., 1923, 123, 822.

² J. Chem. Soc., 1923, 123, 3006.

Norrish and Jones, J. Chem. Soc., 1926, 55.

⁴ Stewart and Edlund, J. A. C. S., 1923, 45, 1014.

Lapworth, Mem. Manchester Phil. Soc., 1920, 64 (iii), 1-16.

A mixed triple bond was also postulated in reactive molecules of acetylene and hydrogen cyanide, HC=-CH and N=-CH, which can then add on the ions of a molecule of water, and thus give rise to the "enolic" forms of acetaldehyde and formamide respectively.

$$\stackrel{-}{\text{HC}}=\stackrel{+}{\text{CH}}+\stackrel{+}{\text{HOH}}\longrightarrow \stackrel{-}{\text{CH}_{1}}\cdot \text{CH}\cdot \text{OH} \longrightarrow \text{CH}_{3}\cdot \text{CHO}$$
 (acetaldehyde)
 $\stackrel{-}{\text{N}}=\stackrel{+}{\text{CH}}+\stackrel{+}{\text{HOH}}\longrightarrow \text{NH}\cdot \text{CH}\cdot \text{OH} \longrightarrow \text{NH}_{3}\cdot \text{CHO}$ (formamide).

Stable Systems Containing Mixed or Semi-Polar Double Bonds.

- "Mixed" double bonds between carbon and carbon or between carbon and oxygen can only be realized as reactive forms in molecules which in their resting state contain non-polar double bonds. The existence of stable systems containing linkages of this kind was first suggested in a paper on "Intramolecular Ionization", which was read before the Faraday Society on November 2nd, 1922, and published early in 1923.¹ The existence of mixed types of valency was then indicated in the following cases:
- (i) In nitrous oxide, O—N—N (compare O—C—O). On the assumption that nitrous oxide and carbon dioxide are isosteric in the sense of containing identical electronic systems around dissimilar nuclei, the change from CO₂ to N₂O implies the transfer of a nuclear proton from oxygen to carbon to make two atoms of nitrogen. The positive and negative charges resulting from this transfer are shown above. In agreement with this formula, E. F. Barker ² has shown that nitrous oxide is an unsymmetrical structure.
 - (ii) In a series of peroxides, e.g.

O=O-O H
$$\stackrel{+}{O}$$
 $\stackrel{-}{O}$ $\stackrel{+}{O}$ $\stackrel{-}{O}$ $\stackrel{C_0H_5}{C_1H_5}$ $\stackrel{+}{N}$ $\stackrel{-}{O}$ $\stackrel{-}{C}$ $\stackrel{+}{C_1H_5}$ $\stackrel{+}{N}$ $\stackrel{-}{O}$ $\stackrel{-}{C}$ $\stackrel{+}{C}$ $\stackrel{+}{C}$ $\stackrel{+}{C}$ $\stackrel{-}{C}$ $\stackrel{+}{C}$ $\stackrel{+}{C}$ $\stackrel{+}{C}$ $\stackrel{-}{C}$ $\stackrel{+}{C}$ \stackrel

¹ Lowry, Trans. Faraday Soc., 1923, 18, 285.

² Phys. Rev., 1931, 38, 1827; Nature, Jan. 23, 1932, 129, 132.

28 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

(iii) In a series of ions, such as

Langmuir 1 had already represented several of these compounds with single bonds between the oxygen and the central atom, but had omitted the electric charges which are needed to complete the scheme. In particular, the representation of univalent oxygen as always carrying a single negative charge (just as quadrivalent nitrogen and tervalent oxygen carry a single positive charge in NH₄⁺ and OH₃⁺) proved rather unexpectedly to be new.

Semi-Polar Bonds Disclosed by the Parachor.

The type of mixed valency which is represented by the symbol + - has been described by Sugden ² as a semi-polar bond. Its actual existence was first established by means of the parachor.³ This function is expressed by the equation

$$[P] = \frac{M}{D - d} \gamma \frac{1}{4}$$

where M is the molecular weight of a substance.

 ${\bf D}$ d are the densities of its liquid and vapour phases. γ is the surface tension of the liquid.

The parachor is virtually a molecular volume corrected for the overwhelming influence of internal pressure as measured by the surface tension of the liquid phase. It has the strictly additive character which was looked for without success in many earlier investigations of molecular refraction, molecular volume, etc. Thus the molecular parachor can be expressed as the sum of a series of atomic parachors, plus a constant for each double or triple bond, or for the formation of a ring of *n*-atoms.

¹ J.A.C.S., 1919, **41**, 868.

² Sugden, Reed and Wilkins, J. Chem. Soc., 1925, 127, 1525.

³ Sugden, J. Chem. Soc., 1924, 125, 1177; "The Parachor and Valency" (Routledge, London), 1930.

TABLE I

Atomic and Structural Parachors

C	= 4.8	Triple bond		46.6
H	= 17·I	Double bond	272	23.2
N	== 12.5	3-membered ring	222	16.7
P	= 37.7	4-membered ring	==	11.6
	= 20.0	5-membered ring	-=	8.5
\mathbf{s}	= 48.2	6-membered ring		6.1
F	= 25.7			
Cl	= 54.3	O ₂ in esters		60.0
Br	68·o			
I	== 01.0			

In the case of compounds which are usually represented with a double bond, however, Sugden found ¹ that (i) whenever the formation of a double covalence is compatible with the octet rule of G. N. Lewis, the molecular parachor is *increased* by 23 units above the sum of the atomic parachors, but (ii) whenever the formation of a double covalence would give rise to a shell of 10 electrons in place of an octet, as in the quinquevalent derivatives of nitrogen and phosphorus, or the quadrivalent or sexavalent derivatives of sulphur, the molecular parachor is *decreased*

TABLE II

Parachors of Non-polar Double Bonds

_ ***	mere of trem perm		22 0	[P] of
	Formula.	[P]obs.	Σ[P] atoms.	Double Bond.
Ethylene	$CH_2 = CH_2$	99.5	78·o	21.5
Propylene	CH ₃ ·CH=CH ₂	139.9	117.0	22.9
Acetone	CH ₃ C==O	161.5	137.0	24.5
Diethyl ketone .	C_2H_5 $C==0$	236.2	215.0	21.2
Carbon disulphide	S=C=S	144.7	101.2	21·8 × 2
Ethyl isothiocyanat	$e C_2 H_3 \cdot N = C = S$	211.7	160.6	25.5 × 2
Nitrosyl chloride .	O=N·Cl	108.1	86.8	21.3
Nitroso- dimethylamine .	CH ₃ N·N=O	184·3	157.2	27.1

¹ Sugden, Reed and Wilkins, J. Chem. Soc., 1925, 127, 1 5.

by about 1½ units for each double bond in the traditional formula for the compound. The compounds with a lower parachor were, therefore, precisely those in which the electronic theory had led to the postulation of a mixed or semi-polar bond.

TABLE III Parachors of Semi-polar Double Bonds [P] of Semipolar Double Bond. Formula. Pots. Σ[P] atoms. Phosphoryl chloride O-PCl, 217.6 220.6 - 3.0 Ethyl phosphate. 399·I 403.0 - 3.9 $OC_{\bullet}H_{\bullet}$ Ethyl sulphite 298.4 299.7 + 1.3 Ethvl ethylsulphonate 295.8 298.4 - 1·3 × 2 Methyl sulphate 238.9 0.75 × 2 240.4 313.8 318.4 196.8 1.65×2 193.3 Sulphuryl chloride Mean

Optical Activity as Evidence of the Dissymmetry of Mixed Double Bonds.

It is a noteworthy rule that each new development of the theory of valency has established itself and proved its utility by leading to the recognition of a new kind of optical activity, as in the cases of (i) van't Hoff's "allene" or centro-asymmetric type of dissymmetry, (ii) the dissymmetric oximes of Mills (verifying the Hantzsch-Werner hypothesis) or (iii) the optically-active co-ordination compounds of Werner. In 1925 a novel type of optical activity was discovered, which could not have been predicted from

the earlier theories of valency, but which found a simple explanation in the theory of "mixed" or "semi-polar" double bonds. This new type of optical activity was detected in the laboratory of the Battersea Polytechnic by Phillips, who was able to resolve into optically-active forms—

- (i) An ethyl toluene-sulphinate, C₇H₇·SO·O·C₂H₅.
- (ii) Two unsymmetrical sulphoxides

CH₃·C₆H₄·SO·C₆H₄·NH₂ and CH₃·SO·C₆H₄·CO·OH.

In the conventional formula the >S=O group is always represented as containing an ordinary double bond. This formula, therefore, gives rise to a symmetrical model, and it is only when the non-polar bond is replaced by a semi-polar

bond, as in the symbol > S-O, that the possibility of optical activity is indicated. The resolution into optically-active forms of a series of compounds of this type, in which the presence of a semi-polar double bond had been established by Sugden's measurements of surface tension and density, was therefore a decisive vindication of the new conception which was embodied in the "polar" formulæ for these compounds.

Dative Bonds.

The shared electrons of a covalence are usually derived equally from the two atoms which they unite, but it is also possible for both electrons to be derived from one atom.² The conditions for this are (a) that the atom which acts as donor should have a lone pair of unshared electrons, which can be shared with another atom and (b) that the atom which acts as acceptor should possess an electronic shell which is capable of assimilating two more shared electrons.

(a) Donors. In the series

H				
н:С:н	н:й:н	н:ö:н	:F:H	:Ne:
H	H	••		

¹ J. Chem. Soc., 1925, 127, 2552; J. Chem. Soc., 1926, 1079.

² Sidgwick f. Chem. Soc. 1923, 123, 728; 1924, 125, 532. Trans. Faraday Soc. 1923, 19, 469.

methane cannot act as a donor, since all the valency electrons are shared between carbon and hydrogen, and, on the other hand, the octet of neon is so stable that sharing of the electrons appears to be ruled out; but ammonia and water are very ready to act as donors, and the same possibility exists in a lesser degree in hydrogen fluoride.

(b) Acceptors. (i) Amongst the acceptors of electrons, the *proton* occupies a foremost place, since it will add on very readily, e.g. to the lone pairs of electrons of ammonia or of water,

$$H: \overset{+}{N}: H + \overset{+}{H} \rightarrow \begin{bmatrix} \overset{+}{H} & \overset{+}{H} \\ \overset{-}{H} & \overset{-}{H} \end{bmatrix}^{+} \qquad H: \overset{-}{O}: H + \overset{+}{H} \rightarrow \begin{bmatrix} \overset{+}{H} & \overset{-}{O}: H \\ \overset{-}{H} & \overset{-}{O}: H \end{bmatrix}^{+}$$

(ii) The tervalent derivatives of boron are also very ready to act as acceptors, since they carry only a sextet of valency electrons

:F: H :F: H
F:B +:N:H
$$\rightarrow$$
F:B:N:H or BF₃ + NH₃ \rightarrow BF₃NH₃
:F: H :F: H

- (iii) The ions of the transition elements (and in general of metals which readily form co-ordination-compounds) also possess the faculty of assimilating electrons; but in most cases this is effected by adding a complete additional shell, usually containing 12 shared electrons. These ions are therefore in a different category from the acceptors cited above, in that they may not be able to accept a single lone-pair of electrons from another atom. On the other hand, the ions of several of the heavy metals (and especially the ions of mercury) are able to assimilate one or two lone-pairs of electrons in their outer shells, since the size of the shell is fixed less rigidly than in the lighter elements. The ionization of mercury salts is therefore generally reversible, like that of the acids, since covalent molecules are formed quite readily.
 - (iv) Finally, we must include the oxygen atom amongst

the acceptors of electrons, since when an oxygen molecule 1 is broken into two atoms, these have only a sextet of elec-

trons : $O:O:D \rightarrow 2:O:$ This is very readily expanded into an octet by accepting lone pairs of electrons from other atoms, as in the oxidation of sulphur dioxide to sulphur trioxide or of a chloride to a perchlorate ion ²

Transfer of Electric Charges in the Formation of Dative Bonds.

When a dative bond is formed, a duplet of electrons belonging to one atom is shared between two. This may be regarded as equivalent to the transfer of a single electron from the donor to the acceptor. If both atoms were originally neutral, the donor then acquires a positive charge and the acceptor a negative charge. Thus when ammonia combines with boron trifluoride

$$BF_3 + NH_3 = F_3 \overrightarrow{B} \cdot \overrightarrow{N}H_3$$

the boron becomes quadricovalent and acquires a negative charge, exactly as it does in the ion $\overline{BF_4}$, whilst the nitrogen in becoming quadricovalent acquires a positive charge, just as it does in the ion $\overline{NH_4}$. The product is a compound containing a semi-polar bond, instead of a covalence. On the other hand, the formation of a dative bond between

The electronic formula assigned here to the oxygen molecule is the one that is usually adopted; but it is obviously inadequate, since it does not account for the paramagnetism of the molecule (see p. 47).

² In general, oxidizing agents accept electrons, as in the case of oxygen, chlorine, the ferric ion, etc., whilst reducing agents give electrons, as in the cases of hydrogen and the ferrous ion $H_{\bullet} = 2H + 2e$. Fe \longrightarrow Fe + e.

34

ions may result in a neutralization of their charges, and the formation of a simple covalence, as in the reversible ionization of mercuric chloride

$$\overline{Cl} + \overline{Hg} + \overline{Cl} \rightleftharpoons Cl - Hg - Cl$$

Finally, the formation of a dative bond between an atom and an ion, as in the cyanhydrin reaction (p. 104), may result in making the ion neutral and the atom charged.

Sidgwick ¹ describes the dative bond as a co-ordinate link, and uses the symbol \rightarrow to represent the direction of transfer of electrons from donor to acceptor. This symbolism is often convenient, as indicating the primary act of forming a compound from a donor and acceptor; but the dative bond, as we have seen, describes a process rather than a product, and confusion may arise if the "co-ordinate link" is identified with a semi-polar bond, since the act of co-ordination may easily give rise to other types of linking. In particular, co-ordination often produces localized charges, which cannot be represented satisfactorily by any system of arrows, since the positive and negative charges do not balance, as in BF₄, NH₄ or SO₄.

The fact that the dative bond represents a process and not a product may also be illustrated by noticing that the use of a separate symbol to portray the act of co-ordination, e.g. in

leads to a discrimination between bonds which are obviously identical, and so revives the difficulty which Werner experienced in deciding whether primary and secondary valencies are different or identical.

The Hydrogen Bond.

- (a) It is an interesting fact that hydrogen, which was long regarded as the immaculate standard of univalency,
- ¹ J. Chem. Soc. 1924, 125, 532, 2672. Trans. Faraday Soc. 1923, 19, 469.

atoms of fluorine, or two atoms of oxygen to one another. Thus, it is clear that the anion of the acid fluorides, e.g. K[HF₂], which has been shown by X-ray analysis to have a linear configuration,1 cannot be held together by the mutual attraction of the two fluoride ions, since each possesses a complete octet and each carries a negative charge; it must therefore be held together by the attraction of two fluorines for the same proton: This attraction can be shown

by writing the formula of the ion as [FHF]; but, as an alternative, it may be supposed that dative bonds are formed between the hydrogen and the halogen ions, giving rise to the complex [FHF], where the bivalent hydrogen is shown with a negative charge.

Similar considerations apply to the structure of ice, in which X-ray analysis has led to the conclusion that every atom of oxygen is surrounded tetrahedrally by four hydrogens, whilst every hydrogen is on a straight line between two oxygens. This structure can be interpreted quite readily on the supposition that dative bonds are formed between contiguous atoms of oxygen and hydrogen. This would give rise to a system in which each quadrivalent oxygen carries a double positive charge, as in the OH, ion, whilst the bivalent hydrogen carries a single negative charge, as in the acid

In this connection it is of interest to notice that, whilst hydrogen can unite two atoms of oxygen or of fluorine, this power is not manifested in regard to chlorine, since there

fluoride ion HF,

is no evidence that the ion HCl, can be formed. Conversely, fluorine differs from the other halogens in the inability of the fluoride ion to unite with four atoms of oxygen to form a perfluorate ion [FO₄]⁻, corresponding with the perchlorate ion, etc., although one or two neutral atoms of fluorine can be linked together by a neutral atom of oxygen in the

¹ Bozorth, J.A.C.S., 1923, 45, 2128.

² Compare Hunter and Samuel, J. Chem. Soc., 1934, 1184.

oxides FO and F—O—F. The origin of these points of difference (which are paralleled by the solubility of AgF and the insolubility of CaF₂) is still obscure.

The theory of bivalent hydrogen was introduced into inorganic chemistry by Moore and Winmill in 1912, when they postulated a link of this kind in order to account for the weak ionization of ammonium hydroxide in contrast with the strong ionization of tetra-methyl-ammonium hydroxide, where the last hydrogen atom of the kation has been replaced by an alkyl group

It was introduced into organic chemistry in 1913 by Pfeiffer,² who explained the feeble acidity of *ortho*-hydroxyanthraquinone by a co-ordination of the hydrogen with a ketonic and a hydroxylic oxygen simultaneously:

The same idea was used by Hantzsch in 1917 3 in order to account for the weakness of many carboxylic acids

$$\begin{array}{cccc} CH_3 \cdot C & & \rightleftharpoons & CH_3 \cdot C & \\ OH & & & & \\ True \ acid. & & Pseudo-acid. \end{array}$$

It was reintroduced into inorganic chemistry by Latimer

- ¹ J. Chem. Soc., 1912, 101, 1635.
- ² Annalen, 1913, 398, 137.
- 3 Ber., 1917, 50, 1422.

and Rodebush ¹ (following Huggins) in 1920, and was applied to a number of additional problems of organic chemistry by Lowry and Burgess in 1923.²

The nature of the hydrogen bond has been a matter of some speculation. As Sidgwick has pointed out, the formation of hydrogen bonds provides a simple explanation of the readiness of compounds containing the hydroxylgroup to undergo molecular association. For this purpose he originally adopted the hypothesis, already set out above, that dative bonds are formed between contiguous atoms of hydrogen and oxygen, giving rise to an aggregate of atoms, which are linked together by a network of covalences, but which must be assumed to acquire electric charges in the process of electron-sharing. Sugden,4 on the other hand, supposed that the atoms are linked by single electron bonds. In his scheme the atoms remain neutral, but the oxygen atoms carry only four shared electrons each in the outer shell, instead of a complete octet.

It is now generally thought that the bivalency of hydrogen is due to "resonance" between two structures $^{\delta}$ (compare pp. 113 and 394) R: H: R \rightleftharpoons R: H: R; for in this way one can, by applying the conceptions of wave-mechanics, explain mathematically the stability of the substances referred to above.

An interesting variant on these earlier views has been put forward by Fowler and Bernal ⁶ in order to account for the "Pseudo-crystalline Structure of Water." They start from the fact that the simple molecule of steam is kinked ⁷ and possesses a strong electrostatic moment. (See Chapter IV,

¹ J.A.C.S., 1920, 42, 1419; compare G. N. Lewis, "Valence and the Structure of Atoms and Molecules", 1923, p. 109.

² J. Chem. Soc., 1923, 123, 2111.

^{3&}quot; Electronic Theory of Valency", 1927, p. 132.

^{4&}quot; The Parachor and Valency", 1930, p. 149.

⁵ Cf. Sidgwick, Chemical Society Annual Reports, 1933, 30, 112-16, 1934, 31, 40-3.

⁶ Trans. Faraday Soc., 1933, 29, 1049.

⁷ At an angle of 105° 6', according to Mecke, Trans. Faraday Soc., 1934, 30, 90.

38

p. 67.) These tripolar molecules can, however, be represented as spheres, since the protons are imbedded in the outer electron-shell of the molecule, thus

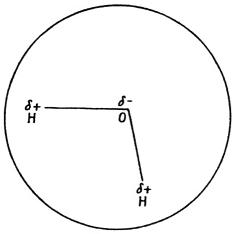


FIGURE II.

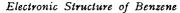
They then conclude that the molecules must arrange themselves in a tetrahedral co-ordination, in which each molecule is surrounded symmetrically by four others; but, instead of placing the hydrogen atoms midway between two oxygen atoms, as in the familiar model of ice, they leave them inside the spheres of the water molecules to which they belong. These spheres are therefore held together by the electrostatic forces between the positive and negative areas on the spheres. The high density of water is explained by supposing that, whilst ice is isomorphous with tridymite, water may be a disordered form of the denser quartz-structure.1 Whilst, however, each proton is placed inside the sphere of an individual molecule of water, it appears that the potential barrier for a proton moving from one water molecule to another is very small. The abnormal mobility of H and OH ions can therefore be attributed to protons jumping

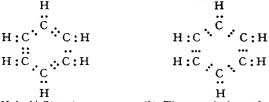
 1 Quartz is stable up to 870° and has a density 2.651 at 0°. Tridymite is stable from 870° to 1470°, and has a density 2.26 at 0°.

from one water molecule to another, either in the direction of the current or against it.

Odd-Electron Bonds.

As soon as it was generally recognized that a bond is formed by two shared electrons, attempts were made to exploit this fact in order to account for the phenomena of residual affinity, etc., by dividing the bond into two equal parts 1 as single electron bonds. An analogous development was the theory of three-electron bonds. This was introduced at a very early stage by J. J. Thomson 2 in order to provide a symmetrical formula for benzene, in which the alternation of single and double bonds was avoided by using two electrons to form each of the bonds between carbon and hydrogen and three electrons to form the bonds between carbon and carbon.





(a) Kekulé Structure.

(b) Thomson's formula.

On similar lines, Löwenbein 3 has proposed to represent the "free radical" pentaphenyl-cyclopentadienyl (c) by the completely-symmetrical electronic formula (d).

$$\begin{array}{c|c}
 & \text{Ph} \\
 & \text{CPh} = \text{CPh} \\
 & \text{CPh} = \text{CPh}
\end{array}$$

$$\begin{array}{c|c}
 & \text{Ph} \\
 & \text{CPh} \\
 & \text{Ph} \\
 & \text{CPh} \\
 & \text{CPh}
\end{array}$$

$$\begin{array}{c|c}
 & \text{Ph} \\
 & \text{CPh} \\
 & \text{CPh} \\
 & \text{CPh}
\end{array}$$

$$\begin{array}{c|c}
 & \text{CPh} \\
 & \text{CPh} \\
 & \text{CPh}
\end{array}$$

¹ Lapworth, J. Chem. Soc., 1922, **121**, 416; Kermack and Robinson, J. Chem. Soc., 1922, **121**, 427.

² Phil. Mag., 1914, [vi] 27, 757.

³ Annalen, 1931, **487**, 97-104. Compare E. Huckel, Trans. Faraday Soc., 1934, **30**, 16.

40 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

The hypothesis of odd-electron bonds finds experimental justification in the existence of a very limited number of compounds, of which NO, NO₂ and ClO₂ are almost the only examples, in which the total number of electrons in the molecule is odd. In addition, there are plenty of examples of "free radicals", where a system containing an odd number of electrons can exist for a short time (like the hydroxyl-radical in steam at 1200° and above), or may even be stabilized by blocking the approach to one another of the unsaturated centres of contiguous molecules. odd-electron compounds are characterized by a magnetic field, which is always developed when the spinning electrons are not united in even numbers to form closed non-magnetic shells. Since none of the ordinary compounds to which oddelectron bonds have been assigned possess magnetic properties, it is clear that these odd-electron bonds, if they exist, must occur in pairs and that their magnetic fields must be so oriented as to neutralize one another. This is possible, but is scarcely likely to be a universal phenomenon.

Modern physics, in providing a firm theoretical foundation for the covalent bond, has not dealt equally kindly with the hypothesis of the single-electron bond, since the principal conditions for binding, namely, the neutralization of the magnetic fields of oppositely-spinning electrons and the exchange-energies of resonating electrons, are both ruled out, when only a single electron is available. There is, however, one complex in which the existence of a singleelectron bond is absolutely certain, namely, the molecular hydrogen ion, H₂⁺, where only a single electron is available to unite the two protons. This complex is easily recognized in a mass-spectrograph; but it is only known in this way, and its transient character may be regarded as a measure of its instability. Borocthane, B.H. is also cited as a possible example of a compound containing a pair of oneelectron bonds. This gas resembles ethane but contains only 16 instead of 18 electrons. If, therefore, the K-electrons of the boron atoms are left intact, there are only 12

electrons with which to form seven bonds. The instability of the gas is regarded as justifying the conclusion that two of the bonds between boron and hydrogen may be formed by one electron only.¹

Sugden's Singlet Linkage.

A much more extensive use of single-electron bonds, under the name of singlet linkages, has been made by Sugden.² Measurements of the parachor have shown a deficiency of about 23 units in the value for phosphorus pentachloride, as compared with the value calculated for

the formula Cl P This deficiency is numerically

equal to that observed in the parachor of phosphorus oxychloride, as compared with the value calculated for the

traditional formula Cl-P-O. In the latter case the

difference is attributed to the formation of a semi-polar bond, PCl_3 —O; in the former case it is attributed to two singlet linkages, as represented by dotted lines in the

formula: Cl₃P(, each producing a deficiency of 11.6

units compared with a two-electron bond.

In these compounds of higher valency Sugden's use of the singlet linkage makes it possible to adhere to the octet rule, without making use of a polar type of formula, such as Langmuir's formula for phosphorus pentachloride [PCl_]Cl. It also provides a convenient method of representing the polyhalide ion in compounds such as KI₃. Thus the linear

¹ Pauling, J.A.C.S., 1931, **53**, 3225.

² "The Parachor and Valency", 1930, chap. VII. Compare Hunter, "The Electronic Theory of Chemistry" (Arnold, London,) 1934, chap. VII.

anion of cæsium-iodochloride 1 can be formulated as Cl I Cl (compare $\overset{-}{F}\overset{+}{H}\overset{-}{F}$); but the sextet of I^+ , like the naked proton H^+ in the formula of the acid fluorides, seems unlikely to represent a final stage in compound-formation; and singlet linkages between the halogens would have the merit of providing three complete octets in the perhalide ions. The conception was also used in 1930 by Hunter and his collaborators in explaining the formation of bromo-addition compounds of heterocyclic bases of the type of pyridine and benzthiazole: 2

On the other hand, if the atoms of hydrogen and oxygen in ice were united only by singlet linkages, the octet rule would be invalid, since each atom of oxygen would be surrounded by only four shared electrons; but the atoms of oxygen and hydrogen would be neutral, as in water, instead of carrying the local charges developed by dative bonds between oxygen and hydrogen. Other illustrations of a similar character have been put forward by Sugden, especially amongst the co-ordination-compounds of diketones; but these need not be referred to here as they are discussed fully in his book on "The Parachor and Valency".

Quadrivalent and Bivalent Carbon.

(a) In the vast majority of its compounds carbon is quadricovalent; but the element is almost unique in that this valency can be developed equally readily in combination with electropositive hydrogen, as in CH₄, or with electronegative chlorine, as in CCl₄. In these compounds four identical bonds are formed by the sharing of four precisely-similar duplets of electrons between carbon and

¹ Wyckoff, J.A.C.S., 1920, **42**, 1100.

² J. Chem. Soc., 1930, 125; J. Indian Chem. Soc., 1931, 8, 147; J. Indian Chem. Soc., 1932, 9, 545; see however: Hunter and Samuel, J. Chem. Soc., 1934, 1183.

hydrogen or chlorine; they may, therefore, be regarded as derived from a form of the element in which the four valency electrons are all in similar orbits. It is designated by spectroscopists as the ⁵S state.

This form of carbon actually exists as an "excited state" of the carbon atom; its existence has been established by the observation of quintuplet lines in the spectrum, pointing to the presence of a system of four uncoupled electrons, since the multiplicity of the lines is one more than the valency of the atom.

This ${}^{5}S$ state of the carbon atoms arises from a system of one s-electron and three p-electrons, and with this configuration the greatest density of the electronic charge distribution occurs at the corner points of a regular tetrahedron. 1

(b) In its ordinary form, denoted as the 3P state, the atom of carbon carries a closed sub-group of two s-electrons and an incomplete sub-group of two p-electrons. In this condition the element might be expected to exhibit a valency of two, since two covalent links might be formed by sharing the two p-electrons with other atoms, without disturbing the more firmly-bound s-electrons. In practice, this lower valency is never exhibited in combination with univalent elements, since CH_2 and CCl_2 are too unstable to persist as permanent molecules; this lower valency is therefore only observed in compounds with bivalent elements, as in CO and CS.

As Langmuir pointed out, 2 carbon monoxide is isosteric with nitrogen, i.e. there is an amazing similarity between the physical properties of the two gases, which he attributes to the fact that the two molecules have identical electronic configurations, although the masses and charges of the nuclei are distributed differently. In the similar case of nitrous oxide and carbon dioxide a symmetrical structure was assigned to carbon dioxide, O—C—O, and a formula for nitrous oxide could then be deduced from it. Carbon monoxide, on the other hand, is a puzzling compound, whereas the properties of nitrogen are expressed very well by the traditional formula, N=N. We may, therefore,

¹ Compare Kronig, "Optical Basis of the Theory of Valency," Cambridge, 1935, page 200.

² J.A.C.S., 1919, **41**, 1543.

now reverse the process and proceed to deduce a formula for the oxide of carbon from that of gaseous nitrogen. This is easily done, since the change consists fundamentally in the transfer of two units of nuclear mass and a single unit of nuclear charge from one atom to another. If no rearrangement of electrons takes place, this leads to the struc-

ture, C = 0, where a semi-polar bond is superposed on a non-polar double bond to form a mixed quadruple bond.

The complete electronic formula is :C:::O:, where each atom carries one "lone-pair" of electrons, and shares three duplets, so that each nucleus is surrounded by a complete octet. This arrangement leaves the carbon atom with a surplus of one electron and the oxygen with a deficit of one electron. A similar structure is assigned to the cyanide

ion: C::: N: (p. 109) and to hydrogen cyanide, H: C::: N: (c) In these systems the carbon is quadrivalent, but one of the four links is an electrovalence, in accordance with the stereochemical considerations (first exemplified by van't Hoff's tetrahedral models) which limit the number of covalent links in a multiple bond to three. Mecke has

covalent links in a multiple bond to three. Mecke has found evidence of the existence of this triple bond in the vibration-spectrum of carbon monoxide, i.e. in the frequency with which the nuclei of the two atoms vibrate to and fro under the influence of the restraining forces set up by the shared electrons. These frequencies are set out for some typical elements and compounds in the following table 1:

TABLE IV

	vioration	Frequencies of	some	Covalent Bonas	
$N \equiv N$	2345	O=O(?)	1565	F-F	1132
$C \equiv O$	2150	$H_3C=O$	1768	H ₃ C—F .	1048
HC==N	2090	$H_2C = CH_2$	1625	H _a C—CH _a	990
HC≡CH	1975			H.NNH.	908
				HO—OH	860

In this table two shared electrons give a frequency of about 1000, four shared electrons give about 1600, and six shared

¹ Trans. Faraday Soc., 1931, 27, 384.

electrons about 2000. The identity of the electronic systems in carbon monoxide and nitrogen, which was postulated by Langmuir as an explanation of the similarity of the physical properties of the two gases, is here confirmed by the similarity of the binding forces of the shared electrons; the two gases, therefore, differ only in the distribution of the nuclear charges and masses.

It may be added that under the influence of ultra-violet light each of these gases gives rise to two "excited" states, one of which is metastable and has therefore a life of appreciable duration. The vibration frequencies of these excited molecules ($N_2 = 1680$ and 1718; CO = 1500 and 1725) correspond with a *double* bond between the two atoms, one linkage of the triple bond being split up by electronic excitation.

(d) Sidgwick and Bowen 1 have deduced the following values for the heats of rupture of covalent links:

TABLE V
Heats of Rupture of Covalent Links

	Abs.*	Rel.		Abs.	Rel.		Abs.	Rel.		Abs. Rel.
C-C	71.0	1	C—N	55	1	C-O	70.5	1	C-S	58·7 I
c = c	125.2	1.8	C=N	111	2.0	c=0	163	2.3	c=s	127 2.2
c = c	165	2.3	—C≡N	183	3.4	c≡o	235.5	3.3		
			-N≡C	184	3.4					

* In Kilogram calories per gram molecule.

"It is remarkable that while with the link of carbon to carbon the heat of formation increases much less rapidly than the number of links, which entirely agrees with what we should expect from the behaviour of unsaturated compounds, with the link of carbon to nitrogen, oxygen, or sulphur the heat of formation of a double link is at least twice, and that of a triple link more than three times, that of a single link."

More recently, Mecke ² and Norrish ³ have suggested that some of these values are subject to a correction for the energy set free when a quadrivalent atom of carbon becomes

¹ Ann. Reports Chem. Soc., 1931, 28, 387.

² Zeit. Electrochem., 1930, 36, 589; Trans. Faraday Soc., 1934, 30, 116, 200.

⁸ Trans. Faraday Soc., 1934, 30, 105.

bivalent. Heitler and Hertzberg 1 in 1929 estimated that the difference of energy between the two forms of carbon, ⁵S-³P was 37 Cal., but the more recent calculations of Van Vleck 2 have lead to the acceptance of the much higher value of 103 Cal. This amount of energy must therefore be allowed for in calculating the energy of dissociation of any process in which a compound of quadrivalent carbon gives rise to atoms of (bivalent) carbon vapour, or to products like CO, in which the carbon is bivalent. Thus the dissociation of methane into atoms of hydrogen and carbon absorbs 374.4 Cal.; but this result includes a liberation of 103 Cal. when the quadrivalent carbon first liberated becomes bivalent. The energy initially required to rupture the four bonds is therefore 374.4 + 103 = 477.4 Cal., and the energy of dissociation of the C-H bond is thereby increased from 93.6 to 122.5 Cal. In this way the observed heats of dissociation deduced by Sidgwick and Bowen 3 must be increased as follows:--

TABLE VI
Heat of Dissociation of Bonds

				Uncorrected.		rrected.
C—H			•	93.6	119.3	kg. cal.
C—C	in	C_3H_6		71·0	122.5	,,
C = C	in	$C_{3}H_{4}$		125.2	228.2	,,
C = C	in	C_2H_2		165	319.5	,,
c=0	in	H ₂ CO		160	211.5	,,
		Me CO		163	214.5	,,
C = C				235.5	235.54	,,
		CH ₃ ·NH ₂		55	81.	,,
		CH ₃ ·NCO		III	162.5	,,
C = N	in	HCN		183	260.	.,
$N \equiv N$	in	N_2 .		-	210.	,,

The Oxygen Bond.

Oxygen has been represented traditionally as containing the same kind of non-polar double bond as ethylene

¹ Zeit. für Physik, 1929, 53, 52.

^{*} J. Chem. Physics., 1933, 1, 177, 219; 1934, 2, 20, 297.

⁸ Ann. Reports Chem. Soc., 1931, 28, 367.

⁴ No correction is needed, because carbon monoxide would dissociate directly to bivalent carbon.

There is, however, a fundamental difference between the two molecules, since oxygen is strongly paramagnetic whilst ethylene is diamagnetic. In terms of the theory of spinning electrons, this means that all the electrons in ethylene are "coupled" to form closed systems, in which the magnetic fields of pairs of oppositely-spinning electrons neutralize one another completely. Quantitative experiments have shown that the magnetic moment of oxygen corresponds with the presence of two uncoupled electrons in each molecule. We can, therefore, assign one uncoupled electron to each atom and represent the molecule by the symbol

where each line represents the two shared electrons of a covalent bond, whilst the two unpaired electrons are represented by broken lines. Theoretical calculations ¹ have shown that in this $^3\Sigma$ state the oxygen molecule has the lowest possible energy level although two of the electrons have unbalanced magnetic fields. This stability of a paramagnetic molecule has been explained by calculations which have shown that in oxygen it is the exchange phenomenon of the two electrons (i.e. the resonance between the molecular orbitals) which plays the dominant rôle, and not the usual spin-coupling.

In addition to its ordinary form, gaseous oxygen is known in two excited states. One of these is metastable and gives rise to a very weak absorption in the red. The vibration frequency (1420) corresponds with a double bond between the two atoms and it is possible that this form of oxygen may be identified as the hypothetical non-magnetic isomer, with two shared duplets, as represented by the conventional formula O=O. The binding is weaker than in the magnetic form, but this may be due to the excitation of the unshared electrons. The other form of oxygen, which gives rise to a very strong absorption in the ultra-violet, has a vibration-frequency of only 708, so that the molecules are now linked only by a single bond. In this highly-activated form of the gas, the electrons of the second duplet appear

¹ Lennard-Jones, Trans. Faraday Soc., 1929, 25, 668; compare Steiner, Trans. Faraday Soc., 1934, 30, 34.

48

to be completely uncoupled; the gas might therefore be expected to be magnetic like atmospheric oxygen.

Nitric oxide contains one electron less than oxygen. Its structure may be derived from that of the non-magnetic isomer of oxygen by removal of one electron from the nitrogen-shell, thus:

The gas is paramagnetic, and quantitative observations show unequivocally that there is one unpaired electron. The only question of importance, then, is to determine whether the odd electron is in an "atomic orbital" belonging exclusively to the nitrogen, or whether it is in a "molecular orbital", and provides part of the binding force between nitrogen and oxygen. Analysis of the band-spectrum shows that the heads of the bands consist of doublets, just like the line-spectrum of sodium; the structure of the molecule, therefore, resembles that of an atom with one valency electron. The vibration-frequency of the molecule (1892) suggests that this "valency electron" is under the control of both nuclei, in the form of a 5-electron bond: N::O: This formula shows a perfectly symmetrical distribution of electrons about the two nuclei, each of

distribution of electrons about the two nuclei, each of which is enveloped by a complete octet; but since it shows three uncoupled electrons, instead of only one, it is perhaps

less plausible than the formula: N::O:, in which the odd electron is associated with the nitrogen, and the nuclear charge of each atom is thus exactly balanced by its electronic shell.¹

In this connection, it is of interest to notice that Kuhn's theory of odd-electron reagents (pp. 128–130) postulates the development of uncoupled electrons in ethylene when acted on by a magnetically-active reagent such as a potassium atom:

¹ See Hunter and Samuel, J. Chem. Soc., 1934, 1181.

It is, therefore, possible that the uncoupled bond of oxygen may have a widespread interest in organic chemistry, in addition to its importance in connection with the formation of peroxides and ozonides by unsaturated substances (p. 150).

The ultra-violet absorption-spectrum of acetylene (beyond 2070 A.U.) discloses the existence of two stable excited states. These may perhaps correspond with the thermal excitation of the gas, where the heat of activation is 34,700 calories per gram molecule, and the photochemical excitation, where the heat of activation is of the order of 100,000 calories.

¹ Bodenstein, Trans. Faraday Soc., 1931, 27, 386.

CHAPTER IV

ELECTRICAL DIPOLES

Symmetrical and Unsymmetrical Molecules.

A neutral atom or molecule is necessarily made up of equivalent positive and negative charges, i.e. the positive charges of the nuclei are exactly balanced by the negative charges of the planetary electrons. In the simple case of an atom of hydrogen, the electron and the proton revolve in unequal orbits round their mass-centre, the orbit of the electron being about 1800 times greater than the orbit of the proton, since its mass is 1800 times smaller; but, when averaged over a period of one or more revolutions, the electric centroids of the positive and negative charges coincide, and the external electrostatic field set up by the opposite charges of the atom is statistically zero.

A similar statement can be made in reference to all other atoms; but in the case of molecules it is the exception rather than the rule for the centroids of the positive and negative charges to coincide. This coincidence, indeed, only occurs in molecules of a high degree of symmetry, such as hydrogen gas, H₂, where the mass-centre of the duplet of shared electrons necessarily coincides with the mass-centre of the two protons. In the same way, the tetrahedral symmetry of methane and of carbon tetrachloride brings the mass-centres of the planetary electrons to coincidence with the mass-centres of the nuclei; and this is also true of the trigonal or hexagonal symmetry of benzene.

At the other extreme are the ion-pairs which constitute the "molecules" of ionized salts such as sodium chloride. These molecules are definitely POLAR, since the sodium ion constitutes the positive pole and the chloride ion the negative pole of a small electrical doublet. The electrical dipole moment of the molecule or ion-pair is defined as the product of the magnitude of the equal and opposite ionic charges and of the distance by which they are separated. Thus, for univalent ions the charges are 4.774×10^{-10} electrostatic units, whilst the distance between the nuclei of sodium and chlorine in a crystal of rock salt is 2.81×10^{-8} cm. A theoretical value for the molecular dipole moment of sodium chloride is therefore $4.774 \times 2.81 \times 10^{-18}$ or 13.4×10^{-18} e.s.u. This moment may be reduced by deformation of the ions 1 or by a displacement of the electrons round the unequally-charged nuclei; but the order of magnitude of the dipole is evidently to be measured in multiples of 10⁻¹⁸ e.s.u.² and has actually been found by the method of molecular beams 3 to be of the order 10-17.

Dipoles may also exist in covalent molecules, provided that these do not possess a centre of symmetry (or an equivalent combination of other elements of symmetry), which would bring the electrical centroids of the nuclei and electrons to coincidence. Thus the symmetrical molecules of hydrogen and of chlorine are both non-polar, but the unsymmetrical molecules of hydrogen chloride are polar, although the two atoms appear to be united by a pair of shared electrons as in the formula H:Cl:. The dipole moments

of these unsymmetrical molecules cannot be predicted, since the separation of the positive and negative centres is not known; but they can be determined experimentally by methods which are described in the present chapter and are found to be of the order of 10⁻¹⁸ e.s.u.

Dipole moments can also be determined for more complex systems, the electrical properties of which can be

¹ Fajans, Z. Krist, 1925, **61**, 29; 1928, **66**, 321; Baker Lectures, Cornell University, 1931, chap. IV, pp. 69-73.

² 10⁻¹⁸ e.s.u. is now known as one *Debye unit* of electrical dipole moment.

Wiede, Zeit. für Physik, 1927, 44, 261; compare Estermann. Zeit. f. physikal. Chem., 1928, B.1, 161.

treated mathematically as if they were composed of opposite charges separated by a definite distance, giving rise to a dipole moment defined by the product of these two quantities. Thus the local electric charges of the betaines, and of compounds containing semi-polar bonds, may be expected to give rise to dipole moments comparable with those of an ion-pair, unless these are cancelled out by being developed symmetrically, e.g. at opposite ends of a benzene ring. Dipole moments may also occur in polyatomic molecules containing only covalent bonds if the distribution of nuclei and electrons is at all unsymmetrical.

Debye's Equation.1

If an electric field could be applied to a perfect vacuum there would be no separation of positive and negative electricity, because there would be no electrons and no positive ions or nuclei in the empty space. When, however, ponderable matter is present, the electrons and the nuclei of the atoms are drawn in opposite directions and the medium is said to be POLARIZED. Under the influence of weak fields, the electrons and nuclei are not torn apart (i.e. the atoms are not "ionized") but are merely displaced relatively to one another. If the medium is regarded as continuous, the POLARIZATION produced by the action of an electric field can be formulated as an electric moment I per unit volume, i.e. a small element of volume dS acquires an electric moment IdS. On the other hand, if the medium is regarded as an aggregate of molecules, we can define its polarizability, a, by the relation

 $m = \alpha \mathbf{F}$

where m is the moment induced in the molecule by a field of intensity, F, and α is the constant ratio of these two quantities.

In order to deal with measurable quantities, we next define the dielectric constant ϵ of the medium as the ratio

¹ Compare Debye, "Polar Molecules", Chemical Catalog Co., New York, 1929; C. P. Smyth, "Dielectric Constant and Molecular Structure", Amer. Chem. Soc. Monograph, New York, 1931. of the electrical displacement **D** to the electrical intensity **E**, in the medium

$$D = \varepsilon E$$

In this equation, the electrical intensity E is measured by the force which would act upon a unit positive charge contained in a long narrow cylindrical cavity in the medium, extending in the direction of the lines of force so that the positive charge is not influenced by induced charges on the walls of the cavity. The electrical displacement, on the other hand, is measured by the force exerted on a unit positive charge when situated in a slot cut perpendicular to the lines of force. It is greater than E, because induced positive and negative charges accumulate on the opposite faces of the slot and increase the force acting on the positive charge. These induced charges depend on the polarization I of the medium in such a way that

$$D = E + 4\pi I.$$

Similarly it can be shown that the force **F** acting upon a unit positive charge in a spherical cavity is given by the equation

$$\mathbf{F} = \mathbf{E} + \frac{4\pi}{3}\mathbf{I}.$$

The polarizability α of the molecules (which is assumed to be independent of the temperature) can be related to the dielectric constant ε by remembering that the polarization I of the medium is the sum of the molecular moments in I c.c. and is therefore equal to mn where n is the number of molecules in I c.c. We can then eliminate I from the equations

$$\mathbf{I} = mn = n\alpha \mathbf{F} = n\alpha \left(\mathbf{E} + \frac{4\pi}{3}\mathbf{I}\right)$$
 and $\mathbf{D} = \mathbf{E} + 4\pi\mathbf{I}$.

On substituting ε for the ratio D/E, we then obtain the equation

$$\frac{\varepsilon-1}{\varepsilon+2}=\frac{4\pi}{3}n\alpha$$

If we now multiply each side of the equation by M/ϱ , where M is the molecular weight and ϱ is the density of the

medium, we obtain the equation of Mosotti (1850) and Clausius (1879), namely

$$\mathbf{P} = \frac{4\pi}{3} \, \mathbf{N} \alpha = \frac{\varepsilon - 1}{\varepsilon + 2} \cdot \frac{\mathbf{M}}{\varrho}.$$

In this equation $N = \frac{Mn}{\rho}$ is the Avogadro number;

 $P = \frac{4\pi}{2}N\alpha$ is a multiple of the polarizability, α , of a single molecule by the Avogadro number and by the factor $4\pi/3$, and is described as the molecular polarization of the substance.

This equation requires correction in all cases in which the molecules have a permanent dipole moment, μ , and are free to rotate in the electric field. This correction is necessary in order to allow for the fact that the molecules will then tend to set themselves along the lines of force, to an extent which will decrease as the temperature increases, and will thus alter the electrical properties of the medium. An appropriate formula to express the influence of this orientation on the electric moment of the medium had already been given by Langevin, who had shown in 1905 that the mean magnetic moment, m, in the direction of a magnetic field F, of gas molecules carrying a permanent magnetic moment, μ , is given by the equation

$$\bar{\mathbf{m}} = \frac{\mu^3}{3KT}\mathbf{F}$$

where T is the absolute temperature and $K = 18.2 \times 10^{-17}$ is the gas-constant for a single molecule. By applying this expression to electric moments, Debye in 1912 and J. J. Thomson in 1914 showed that the molecular polarizability of a medium containing mobile electric dipoles is given by the equation

$$P = \frac{\varepsilon - I}{\varepsilon + 2} \frac{M}{\varrho} = \frac{4\pi}{3} N \left(\alpha_0 + \frac{\mu^2}{3KT} \right)$$

where α_0 depends on the distortion and $\mu^2/3KT$ depends on the orientation of the molecules.

This equation is of the form P = a + b/T so that P is a linear function of r/T. The permanent dipole moment of a molecule can therefore be deduced from the influence of temperature on the dielectric constant by means of the relation $b = \frac{4\pi}{9} \frac{\mu^2 N}{K}$, or, inserting the numerical values of π , N and K,

$$\mu = 0.0127 \sqrt{b} \times 10^{-18}$$

Thus, for ammonia, where a=5.45 c.c. and b=15,250 c.c., $\mu=1.57\times 10^{-18}$ e.s.u.

It should be noticed that the *polarizability* α of the individual molecule and the *molecular polarization* \mathbf{P} of the medium are both expressed as volumes, in agreement with the original assumption of Mosotti that the dielectric constant is a measure of the space occupied by (conducting) particles in the medium. Provided that the molecules are not polar, \mathbf{P} is approximately equal to b/4 in van der Waals' equation (not the b in Debye's equation), which represents the total volume occupied by the molecules, and Debye has calculated that for an atom of hydrogen $\alpha = \frac{9}{2}a^3$, where a is the radius of the orbit of the electron.

It may be added that Debye's equation for the computation of dipole moments may also be derived by the methods of wave-mechanics, and hence is consistent with the modern undulatory theory of the nature of matter.

Detection of Dipoles.

(a) Debye's Method. The most obvious property of a dipole is its tendency to rotate in order to set itself in the direction of the lines of force of an electrostatic field, just as a magnet tends to orient itself in a magnetic field. As the temperature is raised, the molecules of a liquid or gas become more and more agitated, and it becomes less easy to orient them in an electrostatic field. The existence of permanent dipoles was therefore detected by Debye in 1912 1 as a result of an inquiry into the reason why dielectric

¹ Phys. Zeit., 1912, 13, 97.

constants, instead of being independent of temperature, often increase rapidly as the temperature falls. Since the temperature coefficient is a measure of the dipole moment, he was able, by means of the formula developed above, to establish the existence of molecular dipoles in a number of liquid organic compounds, ranging from $\mu = 3.4 \times 10^{-19}$ in methyl alcohol and acetone to 11.8 \times 10⁻¹⁹ in ether. The mean value of the moment in these covalent compounds corresponded with unit charges at a distance 1.1 \times 10⁻⁹ cm., which is 10 times smaller than the typical atomic dimension of about 10⁻⁸ cm.

This is now the most important method for determining the dipole moments of gases and vapours, but the measurements are not easy to make on account of the smallness of their dielectric constants. It cannot be applied to polar liquids, on account of the mutual interaction of contiguous dipoles, so that the moments deduced by Debye in 1912 would not now be accepted as valid; but it is used in a modified form to determine the molecular dipoles of polar compounds dissolved in non-polar solvents, such as benzene or carbon tetrachloride (see below, p. 62).

(b) **Thomson's Method.** The preceding method was described independently in 1914 by J. J. Thomson, who also pointed out that molecules with a permanent dipole moment might be expected to show large deviations from Maxwell's Law (that the square of the refractive index n is equal to the specific inductive capacity ε), since the dipoles would not have time to orient themselves in the rapidly alternating field of visible light. This criterion enabled him to separate gaseous molecules into two classes as follows:

II. H2O NH3 SO, HCI CH2OH C2H5OH CH3CI CHCI3.

The gases of the second series gave evidence of the presence of dipoles, which were absent in the gases of the first series.

¹ Phil. Mag., 1914, 27, 762.

Thomson attributed these dipoles to "intramolecular ionization"; but, since the molecules in question are all covalent, this term must now be interpreted as implying only a partial separation of positive and negative charges, and not the complete ionization that is postulated in the betaines, etc. In a still more approximate form Debye points out that "if ε differs appreciably from n^2 , or what amounts to the same thing, if $\varepsilon - 1$ is appreciably larger than 2(n-1), the molecule is polar". Since, therefore, refractive indices do not often exceed 1.5, a dielectric constant greater than 2 usually indicates that the molecules are polar.

In a more precise form, instead of comparing the dielectric constant with the square of the refractive index, we may consider the molecular polarization P as being made up of two parts, one $P - P_0$ depending on orientation and the other P_0 on distortion. The total polarization P is given by $\frac{\varepsilon - 1}{\varepsilon + 2} \frac{M}{\varrho}$, and the part depending on distortion is given by the analogous expression of Lorentz and Lorenz for molecular refraction $\frac{n^2 - 1}{n^2 + 2} \frac{M}{\varrho}$. Thus for ammonia P = 57.6 c.c. at 292.2° abs., and $P_0 = 5.7$ c.c. so that $P - P_0 = 51.9$ c.c. and

$$\mu = 0.0127 \times 10^{-18} \sqrt{(P - P_0)T} = 1.56 \times 10^{-18}$$

as compared with 1.53×10^{-18} by Debye's method.

Since P is often enormously greater than P_0 (as in the case of ammonia, where it is 10 times greater) the wavelength for which the refractive index is given, is not very important, and the value for the D line of sodium is therefore often selected. Some workers have attempted to avoid this source of error by extrapolating the refractive indices to infinite wave-length by means of a formula of the Selmeier type. The results obtained in this way are, however, still subject to correction by reason of the fact that

¹ Where n is the refractive index of the medium.

the molecular polarization of a medium is made up of three parts

$$P = P_E + P_A + P_{\mu}$$

where $P_{\mathbf{E}}$ depends on the displacement of electrons. $P_{\mathbf{A}}$,, ,, nuclei.

whilst P_{μ} depends on permanent molecular dipoles. Debye has calculated that dipolar molecules would not have time to orient themselves in the direction of an alternating electrical field if the frequency of the field were greater than about 1010 vibrations per second, corresponding with a wave-length of the order of a centimetre. For similar reasons the nuclei, which have a period of vibration corresponding with a wave-length in the infra-red, could not follow the more rapid vibrations of visible light. The refractive indices therefore depend exclusively on the displacement of electrons, as represented by the term P_{r} , whereas the true value of P_{μ} is obtained by deducting $P_{E} + P_{A}$ from the observed polarization P. The magnitude of the error which is introduced by ignoring the "infrared term", PA, can be estimated in some cases by comparing the values of P_{μ} as deduced by methods (a) and (b). In this way K. L. Wolf 1 finds that PA is about 10 to 20% of P_R, although this proportion is probably too high in other cases.2 This error is, however, partially compensated when P_A is deduced from refractive indices for visible light, instead of for infinite wave-length, thus providing a theoretical justification of the convenient practice of using refractive indices for the sodium and mercury lines.

(c) **Ebert's Method.** Orientation of the dipoles is usually impossible in solids. Permanent dipoles may there-

¹ Phys. Zeit., 1930, 31, 227. See Freudenberg's "Stereochemie", p. 216; also C. P. Smyth, "Dielectric Constant and Molecular Structure", New York, 1931.

² See Debye, "Polare Molekeln", Leipzig, 1929; Sutton, Trans. Faraday Soc., 1934, 30, 734; Jenkins, Trans. Faraday Soc., 1934, 30, 739.

⁸ Zeit. phys. Chem., 1925, 113, 1.

fore be disclosed by an abrupt decrease of the dielectric constant on solidification, e.g.

Thus "sulphur monochloride" is shown by these data to have a permanent dipole moment, whilst the molecules of chlorine are non-polar. In terms of polarizations, the following data may be given:

Water (polar) .
$$\mathbf{P}_{\text{LIQ}}$$
 17.8 $\mathbf{P}_{\text{SOLID}}$ 9.5 $(\mathbf{P} - \mathbf{P}_0) = 8.3$
Benzene (non-polar) \mathbf{P}_{LIQ} 26.5 $\mathbf{P}_{\text{SOLID}}$ 25.9 $(\mathbf{P} - \mathbf{P}_0) = 0.6$

In criticism of this method it may be urged that the values of P₀ deduced from the dielectric constant of the solid are not trustworthy, since (i) the Clausius-Mosotti formula only applies to isotropic media and is not applicable to anisotropic crystals and (ii) the electronic shells in the crystal are probably much distorted as compared with those of a liquid or gas; but here again the values for P₀ are often so small compared with the total polarization, P, that minor errors in determining the former are not of any great importance.

A more interesting example of the breakdown of this method is provided by crystals in which certain radicals, or even whole molecules, are able to rotate without disturbing the crystal-lattice. Thus, Kobeko and Kurtshatow have shown that within certain limits of temperature the dielectric constant of crystals of Rochelle salt may rise to values of the order of 20,000! This amazing result may be attributed to the simultaneous orientation of some millions of molecules of water of crystallization under the influence of an electrostatic field. The orientation of these "pivoted" molecules of water of crystallization is much more perfect than anything that can be observed in liquid

¹ Zeit. für Physik, 1930, 66, 192; compare W. G. Cady, Rep. Nat. Res. Council, May 1918; J. A. Anderson, ibid., April 1918; J. Valasek, Phys. Rev., 17, 475, 1921; 19, 478, 1922; 20, 644, 1922; 24, 560, 1924.

water, and so gives rise to enormously larger dielectric constants. Analogous results are to be looked for in other crystals in which polar radicals are free to rotate. In all such cases Ebert's method is obviously invalid, since orientation may actually be facilitated instead of being inhibited altogether by solidification of the medium.

Large values for the dielectric constant of a solid have also been observed by Errera and Ketelaar 1 in anatase and rutile, two polymorphic forms of titanium dioxide. The following data are given

TABLE I Physical Properties of Anatase and Rutile

		Density.	no.	•	$(e - n_0^2)/e$
Anatase		ვ∙85	2.346	48∙0	o·886
Rutile		4.15	2.505	114.0	0.946

In these crystals the "atom polarization" PA is abnormally high, i.e. large moments are developed by deformation of the ions of the crystal at the low frequencies used in measuring dielectric constants, but not at the frequency of light. In the absence of mobile permanent dipoles, the atom polarization P_A amounts to 88.6% of the total atomic and electronic polarization, $P_0 = P_E + P_A$, in anatase and to 94.6% in rutile, and the dielectric constants ε of the two solids amount to 48 and 114, instead of 2 or 3 as in the majority of solids.

(d) Estermann's Method. In a symmetrical magnetic field magnets tend to orient themselves in the direction of the field, but in an unsymmetrical field they tend to move towards the strongest portion of the field. Electrical dipoles tend to act in a similar manner in electrostatic fields.

The method of deflecting a stream of molecules by passage through an unsymmetrical magnetic field has been used with great success by Stern and Gerlach 2 for the investigation of the magnetic properties of atoms and molecules. A similar method, by deflecting a stream of molecules by passage through an unsymmetrical electrostatic

¹ Journal de Physique, 1932, [7] 3, 239.

² Zeit. für Physik, 1921, 8, 110; 1922, 9, 349.

field, has been developed by Estermann 1 for studying the properties of electrical dipoles. Although qualitatively successful, this method has not as yet been developed sufficiently to give results of any quantitative value. The observed deflections are only of the order of 0.1 to 0.5 mm., and so cannot be measured to any great accuracy. As a method of detection of dipoles, however, the method of investigation of dipoles by "molecular rays" is freer from theoretical objection than any of the commoner methods. In fact, it is the only method which affords a direct proof of the existence of electrically dipolar molecules, for all other methods of investigation assume the validity of involved physical theories.²

- (e) Qualitative Methods. There are a number of properties by which polar molecules can be recognized, although they cannot yet be used to calculate the dipole moments of the molecules.
- (i) Hildebrand attributes the wide deviations of binary liquid mixtures from Raoult's Law of Vapour Pressures, leading in extreme cases to the formation of two liquid layers, to the unequal polarity of the liquids. The formation of two liquid layers is therefore a general indication that at least one of the liquids is strongly polar.
- (ii) A molecule with a large dipole moment produces a relatively strong field of force round the molecule. Since these stray fields are largely responsible for the "van der Waals' forces" in a liquid, it is a general rule that substances whose molecules are completely non-polar (or are only weakly dipolar) have a lower boiling-point than those with polar molecules. The low boiling-points of helium,

¹ Zeit. phys. Chem., 1928, **B1**, 161; the earlier application by Wrede (Zeit. für Physik, 1927, 44, 261) of the method of molecular rays to study the dipole moments of the ion-pairs of binary salts (KI, TII, NaI, CsCl and RbBr) has already been referred to.

² See Estermann's paper in "The Dipole Moment and Chemical Structure" (Leipzig lectures, 1929, English translation), pp. 15-21,

Lennard-Jones, Lecture on "Cohesion", Proc. Phys. Soc., 1931, 43, 461-82.

hydrogen and methane, and the high boiling-point of water, can be interpreted in this way. A similar rise of boiling-point may also be expected in liquids such as glycol which contain two dipolar groups, although these may be oriented in such a way that the dipoles neutralize one another.

(iii) Dipole molecules are also often associated with high melting-points and sublimation-points. In this respect, sulphanilic acid and the betaines show some similarity to ionized salts. This phenomenon may, however, be developed also in multipolar molecules, with no total dipole moment. On the other hand, large molecules with only a small local dipole, such as the alcohols, usually melt at quite low temperatures.

Comparison of Results. Consistent values for dipole moments of many substances have been obtained by using methods (a), (b) and (c) above. Each of these has its utility, but the temperature method is probably the most trustworthy. The following table gives a comparison of results obtained by the three methods:

TABLE II
Comparison of Dipole Moments

			P	•			µ X IC)16
Substance.		(a)	(6))	(c)	(a)	(b)	(c)
Phenyl ether .		54.0	52.9		53.0	1.0	1.0	1.0
Benzophenone .		100.0	56.4	51.5	67	2.5	3.0	2.8
Ethyl ether .		31.3	22.5	21.6		1.1	1.25	
Anisole		52.7	32.7	30.0		0.8	1.28	
Aniline		25	30' 3	27.4		1.6	1.5	
Mc-o-aminobenzoate	е.	47.5	42.3		51.2	1.0	1.15	ი∙9ნ
Me-m-aminobenzoat	e	52			56	2.4		2.4
Me-p-aminobenzoat	e.	(40)			56	3.3		3.2
Me-benzoate .		39	37.5		-	1.8	1.8	-

Measurement of Dipoles in Solution.

The preceding theory, by which the magnitude of a permanent dipole moment is deduced from its effect in increasing the dielectric constant of a fluid, depends on the Maxwellian distribution of velocities in a perfect gas. It is therefore not applicable to liquids, nor even to gases under pressures at which the deviations from the gas-laws are

already appreciable. Exact measurements are therefore extremely difficult to make, since the dielectric constants of gases do not differ much from unity. It is therefore necessary to measure these constants to 5 or 6 decimals in order to give the value of the dipole moment to a few units per cent. by means of the approximate expression

$$P = \frac{\varepsilon - I}{3}M$$
, where $(\varepsilon + 2) = 3$ and $\varrho = I$ approximately in Debye's equation (p. 54). Since, however, the gas laws can be applied also to dilute solutions, many investigators have used the much simpler experimental method of measuring the dielectric constants of solutions in order to deduce values of the molecular dipole moment of a

The polarization of a mixture of any two substances is given by the equation:

$$\mathbf{P_{12}} = f_{1}\mathbf{P_{1}} + f_{2}\mathbf{P_{2}} = \frac{f_{1}\mathbf{M_{1}} + f_{2}\mathbf{M_{2}}}{\varrho} \cdot \frac{\varepsilon - \mathbf{I}}{\varepsilon + 2}$$

where f_1 and f_2 are the molar fractions of the two constituents of individual polarizations P_1 and P_2 , and molecular weights M_1 and M_2 . This may be written in the form

$$P_2 = \frac{P_{12} - P_1}{f_2} + P_1$$

for the calculation of P_2 if P_1 is known.

solute.

Sugden ² has proposed an alternative equation, which necessitates a much less involved arithmetical calculation:

$$P_2 = M_2 p_2 = M_2 (p_1 + \frac{p_{12} - p_1}{w})$$

in which p is the specific polarization, $p = \frac{\varepsilon - 1}{\varepsilon + 2} \cdot \frac{1}{\varrho}$

and w is the weight fraction of solute of molecular weight M_2 .

As in the calculation of molecular weights, the value of P_{12} should be determined for several different concentra-

¹ C. P. Smyth, Chemical Reviews, 1929, 6, 549.

² Trans. Faraday Soc., 1934, 30, 720.

64 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

tions, and then extrapolated to infinite dilution. In this way any effects due to association between the molecules of the solute can be eliminated.

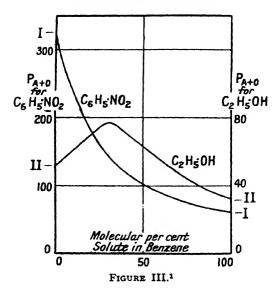
In the normal procedure a non-polar liquid is chosen as solvent. Furthermore, it is assumed that the polar properties of the solute are not altered by any association of a solvent and solute molecules. Suitable solvents for the measurement of dipoles in solution are benzene, cyclohexane, dioxan and carbon tetrachloride, which are all non-polar on account of the symmetrical structure of the molecule. Certain trans-derivatives of ethylene (e.g. trans-CHCl=CHCl) and para-derivatives of benzene might also be used for this purpose, since their dipole moments are zero. As in the determination of molecular weights, the success of the measurements depends upon a correct extrapolation to infinite dilution, and it is therefore necessary to make very exact measurements of the small variations in the dielectric constants of dilute solutions. These can probably be made most easily by a comparison of tuned oscillating circuits 1; but low-frequency bridge methods of balancing capacities have also been found useful.

Influence of Association on Dipole Moments.2

Measurements of polarization in solutions containing high molecular concentrations of polar substances indicate that very large changes may be produced by association between dipoles. The effective polarization of the polar constituent may be either increased or decreased by increase in concentration. Thus in benzene, the polarization of ethyl alcohol is increased whilst that of nitrobenzene is decreased by increase of concentration.

¹ Compare Estermann, "Electrische Momente von Moleculen", pp. 274-9.

² Debye, Polar Molecules, p. 58; C. P. Smyth, Chemical Reviews, 1929, 6, 549-87; Errera, "Molekularassoziation" in Leipziger Vorträge, 1929; Estermann, "Elektrische Dipolmomente von Molekulen", pp. 265-72; J. W. Williams, Chemical Reviews, 1929, 6, 589-619.



This can be explained by supposing that the molecules agglomerate together in different ways, e.g.



Dipole diminished.

Dipole increased.

Association between dipoles is probably small in non-polar solvents, the molecules of which tend to cluster round the individual poles, and thus prevent them from coming into close contact with those of neighbouring molecules. Since the electrical charges of a dipole are usually concentrated upon particular atoms, the degree of association will also depend upon the nearness of these atoms to the surface of the molecule. Since the force exerted by a dipole upon a polarizable molecule is inversely proportional

¹ Reproduced, by permission, from N. V. Sidgwick, "The Covalent Link in Chemistry", p. 172 (G. F. Baker Lectures, Cornell University), Cornell University Press, 1933.

66

to the seventh power of the distance, the radius of effective action of a dipolar molecule will not extend beyond its immediate neighbours, and the molecular movement in the liquid will therefore tend to break up any large agglomeration of dipolar molecules.

Dipole Moments and Chemical Structure.

The dipole moments of all molecules are of the order of 10⁻¹⁸ e.s.u. and though measurements of them are as yet only reliable to a few units per cent. it has been possible to derive much valuable information from them. The data may be examined under the following classification:

- (a) Simple Molecules. Zero values for the dipole moment have been found in
 - (i) Monatomic gases:

Na, K, He, Ne, A, Kr, X.

(ii) Symmetrical diatomic gases of the type AA H₂, N₂, O₂, Cl₂, Br₂, I₂.

(iii) Symmetrical linear triatomic gases of the type BAB, which are non-polar by compensation of opposite dipoles,

(iv) Tetrahedrally-symmetrical molecules of the type AB, namely

CH4, CCl4, SiCl4, SnI4.

On the other hand, substantial dipole moments are observed in

- (i) Unsymmetrical linear molecules of the type AB, e.g. CO o·1, HCl 1·04, HBr o·79, HI o·38 × 10-18 e.s.u. 4
- (ii) Unsymmetrical linear molecules of the type BAC, e.g. O=C=S 0.65, H·C \equiv N 2.6×10^{-18} e.s.u.

but nitrous oxide, which has been shown by spectroscopic evidence 2 to have an unsymmetrical structure, NNO, has no appreciable dipole moment.

- ¹ Lennard-Jones, "Cohesion", Proc. Phys. Soc., 1931, 43, 470. Compare chap. XI.
- ² E. F. Barker, Phys. Rev., 1931, 38, 1827; Nature, Jan. 23, 1932, 129, 132.

- (iii) Non-linear molecules of the type AB₂, e.g. H₂O 1·84, H₂S 0·93, SO₂ 1·61 × 10⁻¹⁸ e.s.u.
- (iv) Molecules of the type AB₃, which are therefore all pyramidal in structure,

NH₃ 1·46, PH₃ 0·55, AsH₃ 0·16 \times 10⁻¹⁸ e.s.u. SbCl₃ 4·0, SbI₃ 0·4

From the point of view of structural chemistry, the existence of dipole moments in molecules such as OH₂ and NH₃ is important, since it suggests that the "lone pairs" of unshared valency electrons in these compounds have a definite orientation in space, just like the shared duplets in fully co-ordinated molecules such as methane, CH₄. The mononuclear orbits of these "lone electrons" must therefore be oriented just as definitely as those of the binuclear orbits of the shared electrons which bind the atoms together. The interchange of shared and unshared electrons, like the interchange of atoms during racemization or as a result of the Walden inversion, probably takes place therefore only by a definite mechanism, for which a critical increment of energy may also be needed.

When the dipole moments of similar compounds in any periodic group are examined, it is seen that, as the atomic weight of an element increases, the polar asymmetry of its compounds becomes less. This seems to indicate that only the arrangement of the outer valency electrons of the atom can produce polar asymmetry, for this will naturally become relatively less in magnitude as the atomic number is increased and successive symmetrical shells of electrons become completed.

(b) Hydrocarbons. (i) Zero values for the dipole moment have been observed, not only in methane and ethane, where the existence of a centre of symmetry effectively prevents the developments of dipole moments, but also in unsymmetrical paraffins containing branched chains, etc. This fact has been interpreted as evidence that the central bond of the system



is electrically symmetrical, even when the hydrocarbon is developed unsymmetrically on either side, and also that the tetrahedral symmetry is preserved.

(ii) The zero dipole moment of CH₄ and CCl₄ is explained by internal compensation of the dipole moments of the four CH or CCl bonds. On this basis, it might be expected that equal dipole moments would be found in

The inequality of the experimental values has been attributed to a distortion of the tetrahedral angle. This has already been established by interferometric methods, which have shown that the distance between adjacent atoms of chlorine is increased from 2.98 A.U. in CCl₄ to 3.11 in CHCl₃ and 3.21 in CH₂Cl₂, corresponding with a spreading of the angle between the CCl bonds from 109° to $116^{\circ} \pm 3^{\circ}$ and $124^{\circ} \pm 6^{\circ}$ respectively.

(iii) Alkyl and Aryl Compounds. Although the symmetrical molecules of methane and benzene have no appreciable dipole moment, a finite value is observed in toluene

$$C_6H_8\cdot CH_8$$
 0.4×10^{-18} e.s.u.

The CH bond which balances the dipole moment of the rest of the molecule in benzene therefore differs from the CH bond which balances the dipole moment of the rest of the molecule in methane, since otherwise these two residues would balance one another in toluene. Analogous differences are observed in the dipole moments of the simple derivatives of methane and benzene, thus:

{CH,Cl C,H,·Cl	1·86 1·55	∫CH ₁ ·NO ₁	3·0 3·8	{CH•COOH	1·73 [1·0]
CH,CN	3.4	C'H'NO	3.90	(C•H•COOH	[0.56]
C.H.CN	3.92	$\begin{cases} (CH_{\bullet})_{\bullet}O \\ C_{\bullet}H_{\bullet}\cdot OMe \end{cases}$	1.32		
{CH•OH	1.67	((CH.).CO	1·23 2·73	CH.NH.	1.23
C'H'OH	1.72	(C.H.),CO (CH.I	2·95 1·65	{c.H.·NH.	1.54
{CH. CHO	2·71 2·75	{C,H,I	1.3	{CH. COOMe	1.74
$\left\{ \begin{smallmatrix} CH_{ullet} Br \\ C_{ullet} H_{ullet} Br \end{smallmatrix} \right\}$	1·80 1·52	C.H.·NO	3.22	C.H. COOMe	1.83

Vector-Addition of Dipole Moments.

From the study of organic compounds it appears that, to a first approximation, a definite dipole moment may be assigned to each individual radical, and that the total dipole moment of the molecule is the resultant of the moments associated with the component radicals, or perhaps with the individual bonds between the atoms.¹ This was recognized in 1914 by J. J. Thomson,2 who recorded the existence of a constant ratio of the dielectric constant to the number of OH, CN, and NO, radicals in unit volume. More recently, it has been shown that the aliphatic alcohols have an almost constant dipole moment, $R \cdot OH = I \cdot 7 \times IO^{-18} (\pm 0 \cdot I)$, whilst the aliphatic ketones have a dipole moment of $R_2CO = 2.7 \times 10^{-18}$ e.s.u. however, the component dipoles are vector quantities, the magnitude of the resultant is determined by the orientation of these moments as well as by their magnitudes.

(a) Cancellation of Dipole Moments in Symmetrical Molecules. The most interesting examples of vectoraddition of dipoles are provided by those molecules in which the component radicals are oriented in such a way that the total dipole moment is zero. This principle has already been invoked in the case of carbon tetrachloride, and can be applied equally well to the hydrocarbons benzene and cyclohexane, where the component dipoles of the six CH radicals in benzene and of the twelve CH bonds in cyclohexane cancel out completely, so that the resultant total moment is zero. Similar conditions apply to many other compounds containing dipolar radicals in symmetrical positions, provided that these substituents are (i) simple radicals, such as the halogens, or (ii) symmetrical radicals, such as methyl. On the other hand, the individual dipole moments in compound radicals may vary in direction as a consequence of the rotation of the atoms about the single bonds, and thus give rise to a resultant total moment in compounds

¹ Eucken and Meyer, Phys. Zeit., 1929, 30, 397.

³ Phil. Mag., 1914, 27, 762.

70 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

which would otherwise be completely symmetrical. Examples of zero moments in symmetrical compounds are found in the following cases:

(i) Trans-derivatives of ethylene.

(ii) Para-derivatives of benzene.

$$X \longrightarrow X CH_3 \longrightarrow CH_3 Cl_3C \longrightarrow CCl_3 NO_2 \longrightarrow NO_2$$
 $(X = F, Cl, Br, I.)$

- (iii) I:3:5-Triderivatives of benzene. $C_6H_3(CH_3)_3$, $C_6H_3(C_2H_3)_3$, $C_6H_3Br_3$.
- (iv) Hexaderivatives of benzene.

 C_6Me_6 , C_6Cl_6 , $C_6Me_3Cl_3$, $C_6Me_3(NO_2)_3$, $C_6Cl_3(NO_2)_3$, $C_3Br_3(NO_2)_3$

(v) I:5 and 2:6 Diderivatives of Naphthalene.

$$CI$$
 NO_{\bullet}

The absence of a dipole moment in 2:6-dichloronaphthalene shows that the two rings are coplanar.

(vi) 4:4'-Di-derivatives of Diphenyl.

where $X = CH_3$, F, Cl, Br, NO_2 .

The absence of a dipole moment in these compounds shows

that the two rings are extended along an axis and not folded over.

(vii). Tetra-derivatives of methane and of tetramethyl methane, not containing bent chains.

$$\begin{array}{ccc} \mathrm{CH_4,} & \mathrm{CCl_4,} & \mathrm{C(NO_2)_4} \\ \mathrm{C(CH_2Cl)_4,} & \mathrm{C(CH_2Br)_4,} & \mathrm{C(CH_2I)_4.} \end{array}$$

- (b) Dipole Moments Depending on Free Rotation.
- (i) The cis and trans derivatives of ethylene of the type

are characterized by a plane and a centre of symmetry respectively, so that only the former can exhibit a total dipole moment. In the corresponding derivatives of ethane of the type $X \cdot CH_2 \cdot CH_2 \cdot X$, free rotation about the central single bond makes it possible for the molecule to assume either configuration in addition to all the intermediate stages. Debye and Ehrhardt 1 by the diffraction of X-rays and Wierl 2 by the diffraction of electrons have shown that the trans position of the chlorine atoms is predominant in the vapour. Since the molecules would have no dipole moment when in this configuration, the existence of a finite moment in compounds of this type proves that other positions are assumed by "free rotation" of the radicals about the central bond. The experimental values are as follows:

(ii) The same phenomenon is observed in penta-erythritol, $C(CH_2 \cdot OH)_4$ and other derivatives of the types CX_4 or $C(CH_2X)_4$, where X is a radical in which the vector of the dipole moment can be displaced by free rotation. Thus, whereas the halides CCl_4 , $C(CH_2Cl)_4$, $C(CH_2Br)_4$ and

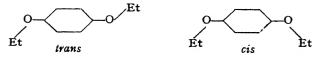
¹ Phys. Zeit., 1930, 31, 142.
² Leipziger Vorträge, 1930.

C(CH₂I), have no moment, the analogous oxygen-derivatives of this group give the following values:

Methyl ortho-carbonate	C(OMe).				0.4
Ethyl ,, ,, .	$C(OEt)_{4}$				1.1
Penta-erythritol	$C(CH_2OH)_4$.				2
Penta-erythritol acetate	$C(CH_2 \cdot OAc)_4$.				2.18
Penta-erythritol nitrate	$C(CH_2 \cdot ONO_2)_4$.				2.0
Methyl methane-	C(CH, CO-OMe)				2.8
tetracarboxylate		-			
Ethyl methane-	C(CH2·CO·OEt)				3.0
tetracarboxylate					
Di-isopropylidene	O·CH ₂	CH.	·O、		
penta-erythritol	Me ₂ C	c⁄ ¯	>c	Me.	2.3
	O·CH ₂	∨CH,	·0/	_	_

These observations are of interest in relation to X-ray analyses of the crystals,1 which have shown that the full tetrahedral symmetry of CCl4, SnI4, etc., is not maintained in these oxygen-compounds.

(iii) Other important examples of dipole moments resulting from free rotation about the single bonds of compound radicals are found in para-di-derivatives of benzene, containing the radicals OH, OCH₃, NH₂, NHMe, etc., where the bonds which link the atoms of oxygen or nitrogen to the benzene ring are inclined at an angle (of the order of 110°) to the bonds which link them to the other atoms of the substituent radical. The "classical" example is found in the diethyl ether of quinol, where free rotation about the single bonds between carbon and oxygen may lead to the existence of a trans phase, with a centre of symmetry and a cis phase with a plane of symmetry



The former would have no dipole moment, since the constituent dipoles are equal in magnitude but opposite in direction. The latter would have a resultant dipole which could be calculated from the individual dipoles by the

¹ See a report by H. Mark in Freudenberg's "Stereochemie", p. 93.

usual parallelogram law for compounding vectorial quantities. The actual existence of dipoles in compounds of this type is shown in the following table:

	Quinol diethyl ether Quinol diacetate . Methyl terephthalate Ethyl terephthalate.	:	EtO·C ₄ H ₄ ·OEt . AcO·C ₄ H ₄ ·OAc . MeO·CO·C ₄ H ₄ ·CO·OMe EtO·CO·C ₂ H ₄ ·CO·OEt	:	:	1·7 × 2·2 2·3	10-11
	Benzidine		NH, C, H, C, H, NH,			1.3 ×	10-18
	Tetramethylbenzidine		NMe.C.H.C.H.NMe.			1.25	
4:4'	Dimethoxydiphenyl.		MeO·C.H.·C.H.·OMe			1.5	
4:4'	Diethoxydiphenyl .		EtO-C.HC.HOEt			1.9	,,
4:4'	Diacetoxydiphenyl .		CH.·CO·O·C.H.·C.H.·O	·CO·CI	н,	1.9	,,

(iv) Similar dipole moments have been observed in the following symmetrical I:3:5-tri-derivatives of benzene, where again the vectors of the component moments need not be in the plane of the ring

(v) In marked contrast to the above examples, paradinitro-compounds have no resultant dipole moment (see p. 70), indicating that the dipole of the nitro group is vectorially directed symmetrically between the two oxygen

atoms,
$$-N \xrightarrow{O}$$
 and not at an angle, as one would

anticipate from the "semi-polar bond" formula
$$-\stackrel{+}{N} \stackrel{O}{\stackrel{-}{\bigcirc}}$$
.

This result, indicating the equivalence of the two oxygen atoms, receives further discussion in a later chapter (p. 396).

Calculation of Dipole Moments by Vectorial Addition.

In order to calculate the total dipole moment of a molecule it is necessary to know (i) the sign, (ii) the magnitude, (iii) the direction of the compound dipoles.

(a) **Sign of Dipole Moments.** From investigations of a single polar substance only the magnitude of the moment, i.e. the product of pole strength and distance, and not its sign, can be determined. In order to ascertain at which

end of the molecule the positive and negative electrical poles are situated, it is necessary to examine a series of related compounds, containing different polar groups in various steric relationships to one another. Thus the sense of the polarization of the groups can be deduced from a study of the dipole moments of the following aromatic compounds:

```
Nitrobenzene . . 3.9 o-Nitro-chlorobenzene 4.3 to 3.8 Chlorobenzene . . 1.55 m-Nitro-chlorobenzene 3.4 p-Nitro-chlorobenzene 2.6
```

The polarizations due to the nitro and chloro groups obviously oppose each other in the para position, and consequently are of similar sign when viewed from the centre of the aromatic ring. In the nitrotoluenes and nitroanilines, however, the individual polarizations of the two polar groups increase the polarity of the whole molecule in the para derivatives. They therefore point in the same direction in space and consequently have opposite senses with respect to the benzene ring.

Toluene	•	•	3·9 0·4	Nitrobenzene Aniline.	•	•	•	3.9
	•	•	0.4		•		•	1.0
o-Nitrotoluene			3.75	o-Nitraniline				4:5
m-Nitrotoluene			4.2	m-Nitraniline				4.7
p-Nitrotoluene			4.5	p-Nitraniline				7.1
(Williams, J.A.C.S.	., 1928	, 50, 36	52.)	(Höjendahl, Phys.	Zeit.	, 1929, 3 (), 3 91.)

The dipole moment of a nitro-compound is undoubtedly due mainly to the semi-polar bond of the nitro group,

$$-\overset{+}{N}\overset{O}{\underset{O}{\left(\right)}}$$
, and consequently is in the sense, $C_{\mathfrak{o}}H_{\mathfrak{o}}\overset{+}{N}\overset{+}{O}_{\mathfrak{o}}$,

where the arrow points from + to -. Hence in chlorobenzene, toluene and aniline the dipoles must be in the senses:

This is exactly what would have been anticipated from general conceptions of atomic properties, since chlorine with its large nuclear charge would be expected to attract electrons in the direction shown by the symbol C—Cl, whilst the protons in the methyl radical represent a concentration of positive charges in the sense of the symbol —CH₃.

The symbol +— is due to Sidgwick. It is used in such a way that the arrow points in the direction leading from the positive to the negative pole of the dipole. It is customary to describe the dipole moment of a group R as positive or negative according as it is directed in the sense shown by the symbols

This convention is easy to remember in view of the fact that the "tail" of the arrow is in the form of a + sign.

(b) Magnitude of Dipole Moments. By dissecting the observed dipole moments of a large number of organic compounds, the following values have been deduced for the component moments of the principal radicals that constitute organic molecules:

TABLE III

Dipole Moments of Individual Links 1 (in 10 18 e.s. units)

HC	0.2	CC	0.00	P-Cl	o·8
H-N	1.5	CN	0.4	P—Br	0.6
HO	1.6	CEN	3.3	AsCl	2.0
HS	0⋅8	C-O	0.0	As—Br	1.7
H-P	0.55	C = O	2.5	SbCl	3.9
HAs	0.15	CEE.O	5.3	N=0	1.9
H-F	(2)	C—F	1.5	CS	1.3
HCl	1.03	CCl	1.7	C==:S	3.0
H—Br	0∙78	C-Br	1.6	C—Se	1 · 1
H—I	o·38	I	1.4	C-Te	0.9

In calculating the dipole moments of the individual links the assumption is made that the actual moment of a link is not seriously affected by the other links that its constituent atoms may form. This assumption is not quite true, for dipole moments are not strictly additive vector

¹ From Sidgwick, "The Covalent Link in Chemistry", Baker Lectures, Cornell University, 1933, p. 153.

quantities, as we have seen already in comparing the dipole moments of alkyl and aryl compounds (p. 68). One can ascribe most discrepancies to the effects of "electrostatic induction" between the different parts of a complex molecule. The induction of electrical dissymmetry will be considered in a subsequent chapter (Chapter XI) and it will suffice to mention here that, apart from specific cases now well recognized, amongst which the difference between aliphatic and aromatic compounds is by far the most pronounced, the effect of electrostatic induction in altering the magnitude of a dipole moment is very small, and frequently is beyond the limits of experimental error.

For all practical purposes it is justifiable to assume that for any specific linkage of atoms there exists one inherent dipole moment.

(c) Direction of Dipole Moments. In accordance with general chemical theory the component dipole moments of substituted derivatives of benzene radiate from the centre of a flat ring, and are therefore inclined at 60°, 120°, or 180° to one another. J. J. Thomson therefore pointed out in 1923 2 that the resultant moment due to individual moments μ_1 and μ_2 in the di-derivatives of benzene can be calculated by vectorial addition by means of the following formulæ:

$$\mu_{ortho} = \sqrt{\mu_1^2 + \mu_2^2 \pm 2\mu_1\mu_2 \cos 60^\circ}$$

$$\mu_{meta} = \sqrt{\mu_1^2 + \mu_2^2 \pm 2\mu_1\mu_2 \cos 120^\circ}$$

$$\mu_{para} = \mu_1 \pm \mu_2$$

(i) In the special case in which the two substituents are identical $(\mu_1 = \mu_2)$ these formulæ become simplified as follows:

$$\mu_{ortho} = \sqrt{3\mu}$$
 $\mu_{meta} = \mu$
 $\mu_{para} = 0$.

The para-di-derivatives, therefore, have a zero moment as

¹ Phil. Mag., 1923, 46, 513.

¹ Compare, inter alia, C. K. Ingold, Chemical Reviews, 1934, 15, pp. 225-74. See also p. 395.

described above. On the other hand, the meta-di-derivatives have the same dipole moments as the mono-derivatives,

e.g.
$$m-C_0H_4Me_2 = 0.40$$
 (0.40)
 $m-C_0H_4Cl_2 = 1.48$ (1.55)
 $m-C_0H_4Br_2 = 1.50$ (1.52)

where the values in brackets are those for the mono-derivatives.

(ii) An instructive illustration of the use of the formula for p-derivatives is provided by the following examples:

CH ₃ ·Cl				1.86	$CH_{s} \cdot Br$			1·80
C ₆ H ₈ ·Cl				1.55	$C_{6}H_{5}\cdot\mathbf{Br}$			1.52
CH ₃ ·C ₆ H ₅		•	•	0.4	CH₃·C ₆ H ₅		•	0.4
$CH_3 \cdot C_6H_4 \cdot C$	1	•	•	1.90	CH ₃ ·C ₆ H ₄ ·B	r		1.93

Not only is the total moment of the p-di-derivative equal to the sum of the component moments of the mono-derivatives, but it agrees closely with that of the methyl halide in which the aromatic ring between the methyl and halogen has been suppressed. Thus, in accordance with the theory of vector-addition, the total dipole moment is not influenced by the distance between the component moments (e.g. \leftarrow + \rightarrow CH₃ and CCl) which are being compounded.

(iii) The values of the total moment calculated from Thomson's formulæ agree closely with the observed values for meta- and para-derivatives, from which we may conclude that the mutual influence of the dipoles is negligible. In the ortho-derivatives, however, there appears to be a spreading of the valency directions between similar dipoles, and perhaps a closing up of dipoles of opposite sign, since the observed values are smaller than the calculated values in the former case, but may be slightly larger in the latter case :

similar sign		•		Obs. 0·55 2·25 2·00	Calc. 0·71 2·67 2·63	0-C + 0·16 + 0·42 + 0·63
Dipoles of opposite sign $\begin{cases} o\text{-CH}_3\text{-C}_6H_4\text{-C}_6\\ o\text{-CH}_3\text{-C}_6H_4\text{-C}_6\\ o\text{-CH}_3\text{-C}_6H_4\text{-C}_6\end{cases}$	Br		•	1·35 1·44 3·77	1·30 1·34 3·74	+ 0.04 - 0.10 - 0.03

(iv) In calculating the dipole moments of oxygenated derivatives of benzene, it is necessary to take account of the free rotation of radicals containing this element. If it be assumed that perfectly free rotation is possible about all the single bonds in a molecule, the resultant average value of the dipole moment will then be the root mean square of the actual values in all possible steric positions. This can rarely be computed, but it has been applied in a simplified form to molecules containing two similar mobile polar groups, in which the actual value could vary from zero to twice the moment of any one of the dipolar groups. For such compounds the root mean square of the dipole moment is $\sqrt{2}$ times the moment of a single group. The para-disubstitution derivatives of benzene are compounds of this category and their polarizations are given below:

Substance			Cal: $(\sqrt{2})$	Observed.
Anisole				1.2 0.8
Phenetole				I·O (± 0·2 ?)
				1⋅8
Quinol diethyl ether			1.4 to 1.6	1.7
Dimethyl terephthalate			2.3	2.2
4:4'-Dimethoxydiphenyl			1.1 to 1.6	1.5
4:4'-Diethoxydiphenyl	•	•	1.4 to 1.6	1.9

Only a rough correlation holds between observed and calculated values.

Valency Deflection and Free Rotation. (a) Although the homologous alcohols have a constant dipole moment $\mu = 1.7 \times 10^{-18}$ e.s.u., this is no longer true of the ethers.

Since the length of the bonds between carbon and oxygen is not likely to vary more in the ethers than in the alcohols, the progressive decrease in the dipole moment, as the size of the alkyl radical increases, is attributed to a spreading of the bonds, similar to that which has been shown to occur in CH₂Cl₂ and CHCl₃ when compared with CCl₄

¹ Ebert, "Dipolmoment und innere Molekulbewegung", Leipziger Vorträge, 1929, p. 64.

- (p. 68). A similar effect has also been indicated in the ortho di-derivatives of benzene (p. 77).
- (b) In the case of oxygenated ortho-derivatives, still larger deviations are observed. These are explained by supposing that in the m- and p-derivatives (where a good agreement with values deduced from the theory of free rotation is observed) all configurations are equally probable whereas in the o-derivatives free rotation is impeded by the close contiguity of the substituents. Since free rotation is likely to give place to preferred configurations as the temperature is reduced, this phenomenon may also manifest itself by a dependence of the permanent dipole moment on temperature within the range in which free rotation is being developed or suppressed. Thus, the dipole moment of hydroquinone diethyl ether is independent of temperature, because the radicals are too far apart to influence one another, whilst that of the carboxyl-group may be constant because the carbonyl and hydroxyl radicals are so close that free rotation is completely inhibited. On the other hand, the dependence of the dipole moment of dichloroethane on temperature, between -70° and $+270^{\circ}$ C., can be interpreted as due to a progressive development of free rotation, in agreement with conclusions reached from the absolute magnitude of the moments.

Dipole Moments of Oximes.

It has already been shown (p. 70) that trans-compounds of the type CHX:CHX, which possess a centre of symmetry, have no dipole moment, whereas dipoles of substantial magnitude are developed in the cis-isomers, which have a plane but no centre of symmetry. This method of investigation has been extended to other than ethylenic compounds. Thus, from measurement of dipole moments, Sutton 2 has provided a clear answer to the much-disputed question of the configuration of the α and β forms of the

¹ L. Meyer, Zeit. phys. Chem., 1931, B8, 27; C. T. Zahn, Phys. Rev., 1931, 38, 521.

² J. Chem. Soc., 1931, 2190.

oximes, and the manner in which they undergo the Beckmann transformation. The N-ethers of p-nitrobenzophenoneoxime were chosen for examination since in the syn-compound (A) the two semi-polar N-O groups should produce large moments in the same sense, but in the antiform (B) these should be in opposite senses. Of the two oximes, the α-form (the high melting isomer) had the anti-configuration, and gave by the Beckmann rearrangement p-nitrobenzoanilide, whilst the β -form gave benzo-p-nitroanilide.

This measurement has provided conclusive physical evidence in support of the theory of Meisenheimer 1 that the Beckmann rearrangement depends on an interchange of groups in the trans position (see pp. 352-354), which had previously been advanced on purely chemical grounds. It leads, however, to configurations which are the reverse of those originally suggested by Werner.

Evidence based upon dipole moments also confirms the assignation of a trans-configuration to the aromatic azocompounds. Thus azobenzene, of which the trans-structure would have a centre of symmetry, has no dipole moment, whilst p-chloro and p-bromo azobenzenes have practically the same dipole moment as chloro- and bromobenzenes.

CHAPTER V

CHEMICAL REACTIVITY

Inorganic and Organic Reactions.

One of the most striking contrasts between organic and inorganic chemistry is to be found in the instantaneous character of most inorganic reactions, such as the double decomposition of pairs of salts, and the tedious character of most of the reactions of organic chemistry, even when the reagents used are inorganic. Thus, whereas the formation of a salt from an acid and an alkali is complete as soon as the solutions have had time to mix, the preparation of an ester from an alcohol and an acid, or the nitration of an aromatic hydrocarbon, generally occupies a period of several hours.

This contrast can be explained in part by the fact that inorganic chemistry is in main the chemistry of ions, whilst organic chemistry is the chemistry of bonds (p. 7). In an interaction between ions the whole of the material is in a condition to undergo chemical change immediately, but when the action involves the rupture of a bond it seems to be necessary to allow time for the molecules to be brought into a condition to undergo chemical change.

Activation.

The conception of activation was introduced by Arrhenius ¹ in 1889 in order to explain why chemical changes are not all instantaneous, like the interaction of a pair of ions, and also why the temperature coefficients of chemical change are so large, since an average doubling of the velocity of chemical change by a rise of temperature of 10° cannot be

G

¹ Zeit. phys. Chem., 1889, 4, 226.

explained by an increase of about 0.3% in the number of molecular collisions per second. He therefore suggested that many chemical changes depend on the presence of a small proportion of active molecules amongst a larger proportion of normal molecules which are chemically inactive in the resting state. Activation of the normal molecules can be brought about by raising them to a higher energy-level. This can be effected most easily by raising the temperature, and thus increasing the kinetic energy of the molecules, or by the action of light, when the absorption of radiant energy raises the more loosely-bound electrons to a higher "stationary state". Catalysts are, however, even more effective than heat and light in promoting chemical change, although the way in which they act is much more obscure.

Energy of Activation.

By the application of thermodynamics to the theory of activation of chemical change, Arrhenius derived the important equation

$$k = B \cdot e^{-\frac{E}{RT}}$$

connecting statistically the velocity, **k**, of any reaction with the energy of activation, or critical energy increment, **E**, necessary to bring about the reaction. This equation, and others that have been derived from it for the purpose of chemical calculation, can only be applied to chemical reactions in bulk, when probability relationships hold true: it cannot be applied to the reactions of isolated molecules, for which the methods of quantum mechanics alone are valid.

Reaction velocities cannot be correlated directly with energy changes alone, since the equation connecting reaction velocity with temperature contains two variable factors, namely **E** the energy of activation, and **B** the reaction constant. The reaction constant depends primarily upon the frequency of collision of the reacting molecules. In homogeneous bimolecular gas reactions it approaches the value

for the frequency of collisions as calculated by means of the kinetic theory.¹ The same statement is true for some reactions in solution,² but in others it is several powers of ten smaller than this.

From a detailed study of the kinetics of chemical changes in solutions Moelwyn-Hughes,³ in 1932, concluded that reactions could be divided into three groups as follows:

(I). Reactions between two ions, leading to the formation of an Ion-pair.

These are instantaneous: in other words E = 0.

Typical examples are $Ag^+ + Cl^+ = AgCl$; and also $H^+ + OH^- = H_2O$, though in the latter reaction the product is a covalent molecule and not an ion-pair.

(II). Reactions between a Covalent molecule and an Ion.

For these reactions the factor **B** in the equation $\mathbf{k} = \mathbf{B}\mathbf{e}^{\mathbf{\overline{RT}}}$ is of the same magnitude as **Z**, the frequency of collision between the reactants.⁴ At concentrations of one grammolecule per litre the factor **Z** has an average value of 2.77×10^{11} for most organic molecules.

Typical reactions of this class are:

$$C_6H_5O$$
 Na + C_8H_6 —I = C_6H_6O — C_8H_6 + Na I
 CH_3CI ·COOH + OH = $CH_8(OH)$ ·COOH + CI

(III). Reactions between two Covalent molecules.

The velocity of interaction of two covalent molecules is many times smaller than that calculated by the equation,

 $k = Ze^{-E}$. Not only does one find that **B** is very much

¹ Hinshelwood, "Kinetics of Chemical Change in Gaseous Systems", 1929, p. 54.

Moelwyn-Hughes, Chemical Reviews, 1932, 10, 241; Traill, Phil. Mag., 1932, [7], 13, 225; Moelwyn-Hughes, ibid., 14, 112; Moelwyn-Hughes, "The Kinetics of Reactions in Solution", Oxford Press, 1933.

[&]quot;Kinetics of Reactions in Solution", Oxford Press, 1933.

⁴ For calculation of **Z**, see Moelwyn-Hughes (loc. cit.); also Bradley, J. Chem. Soc., 1934, 1910.

smaller than **Z**, but also **E** is not exceptionally large. The slow velocity cannot be attributed entirely to an effect of the solvent, but depends upon the process by which the chemical change is brought about, and much further investigation of this subject is still needed.¹

Typical slow reactions of this class are:

$$C_5H_5N + CH_3I = C_5H_5N - CH_3I$$
 (pyridine methiodide)
 $C_6H_5NH_2 + C_6H_5 \cdot CO \cdot Cl = C_6H_5 \cdot NH \cdot CO \cdot C_6H_5 + HCl$

Ionic and Non-Ionic Reactions.

According to the electronic theory of valency, there are two ways in which atoms united by a covalent bond can be separated, namely (i) by a symmetrical dissociation into neutral radicals, with a severance of the two electrons of the shared duplet, or (ii) by an unsymmetrical dissociation into ions, whereby the duplet is transferred as a whole to one or other of the atoms which it formerly united. The first method may be termed atomic reaction, since it corresponds to dissociation of a molecule into its constituent atoms with their characteristic electron groupings; or, since it does not lead to the production of any electrical charges, it may alternatively be designated as non-ionic reaction. The second method of reaction, depending on dissociation and recombination of ions, may, by contrast, be termed ionic reaction.

Since there is no other method of disposing of the two electrons of a broken bond, all organic reactions must be assigned to one or other of these two types of chemical change, although in practice "border-line" cases may be encountered in which a correct classification cannot easily be made on the basis of existing experimental knowledge.

¹ Hinshelwood, J. Chem. Soc., 1933 1357; 1935, 1111; 1936, 371; 1936, 1353.

Moelwyn-Hughes, J. Chem. Soc., 1936, 101: Acta Physicochimica U.R.S.S., 1936, 4, 173-224. Chemical Society, Annual Reports, 1936, pp. 106-109.

Hughes and Ingold, J. Chem. Soc., 1935, 244; 1936, 225.

Non-Ionic Reactions.

In the following paragraphs the view is developed that the two possible types of chemical reactions occur under characteristic experimental conditions and proceed in distinctive ways which can be differentiated from each other.¹ Thus the characteristics of non-ionic reactions may be illustrated by the following examples.

- (a) Photochemical Changes. The photochemical combination of hydrogen and chlorine is now believed to take place through the intervention of neutral atoms, since attempts to detect electrical conductivity in the reacting gases have failed.² The separation of the halogen molecule into two reactive atoms, instead of into a reactive positive and a non-reactive negative ion (p. 184), is disclosed experimentally in this and in other reactions, by the fact that the rate is proportional to the square root of the concentration of the halogen.³
- (b) Reaction Chains. In accordance with Einstein's law of photochemical equivalence, each quantum of light which is absorbed by a photosensitive substance activates a single molecule, so that the number of quanta absorbed and the number of molecules decomposed are usually equal to one another; but in other cases the absorption of a quantum of light may bring about a chemical change in some millions of molecules. It has therefore been suggested that, in cases of the latter type, the products of the initial photochemical decomposition are able to propagate a long reaction-chain. Many gaseous explosions are believed to be propagated in this way.

Chain mechanisms, involving ionized atoms, were first suggested by Bodenstein in 1913 4; but, since the formation

¹ Waters, Proc. Durham Phil. Soc., 1932, 8, 459; Lowry, Trans. Faraday Soc., 1934, 30, 3.

² J. J. Thomson, *Proc. Camb. Phil. Soc.*, 1901, 11, 90; Ludlam, *Phil. Mag.*, 1912, 23, 757; Le Blanc and Volmer, *Zeit. Elektrochem.*, 1914, 20, 494.

³ Bodenstein and Lind, Zeit. phys. Chem., 1906, 57, 167.

⁴ Zeit, phys. Chem., 1913, 85, 329.

of free ions was shown to require the consumption of too much energy, they were abandoned, first for mechanisms involving activated molecules 1 and subsequently for mechanisms postulating the dissociation of the halogens into atoms, 2 e.g.

$$\begin{array}{l} \operatorname{Cl}_{2} + \operatorname{H} \longrightarrow_{l} 2\operatorname{Cl} \\ \operatorname{Cl} + \operatorname{H}_{2} \longrightarrow \operatorname{HCl} + \operatorname{H} \\ \operatorname{H} + \operatorname{Cl}_{2} \longrightarrow \operatorname{HCl} + \operatorname{Cl} \end{array} \right\} et \ seq.$$

More recently it has been shown 3 that the combination of hydrogen and chlorine can be initiated by atoms of chlorine, formed by the action of sodium-vapour on molecules of chlorine in a mixture of hydrogen and chlorine. 4 The chain reaction is then formulated as follows:

$$\begin{aligned} \text{Na} &+ \text{Cl}_2 &= \overset{+}{\text{NaCl}} + \text{Cl} \\ \text{Cl} &+ \text{H}_2 &= \text{HCl} + \text{H} \\ \text{H} &+ \text{Cl}_2 &= \text{HCl} + \text{Cl} \end{aligned} \right\} et \ seq.$$

Other gaseous reactions which do not depend on the formation of ions include the oxidation of hydrogen, carbon monoxide, methane, and other hydrocarbons. These oxidations can be initiated photochemically 5 in presence of a trace of ammonia. They then proceed either at moderate temperatures and measurable rates, or at high temperatures by means of an explosion wave which spreads through the reaction mixture. These oxidations are probably initiated by hydrogen atoms from a primary photochemical decomposition of ammonia and have all the characteristics of chain-reactions.

(c) **Predissociation.** The decomposition of an activated molecule obviously depends upon the internal distribution of the available energy, which has to find its way to the critical linkage at which decomposition takes place.

¹ Zeit. Elektrochem., 1916, 22, 58.

² Nernst, Zeit. Elektrochem., 1916, 22, 62; 1918, 24, 1916; Warburg, Sitzungsber, Preuss. Akad., 1916, 314; 1918, 300.

⁸ Polanyi, and others, Zeit. phys. Chem., 1928, B1, 3 et seq.

Franck, Zeit. für Physik, 1922, 9, 259; Faraday Society Discussion on Photochemistry, 1925, 21, 536.

⁵ Haber, Farkas and Harteck, Zeit. Elektrochem., 36, 711 (1930).

If this distribution is unfavourable, the active molecule may survive without decomposition during an appreciable period, even when it has absorbed more photochemical energy than is required for simple thermal dissociation. The onset of photochemical reactivity is often revealed by a blurring of the fine line-structure of the absorption bands. which has been described by its discoverer, V. Henri, as predissociation. The blurring, which takes place when the wave-length of the light is reduced below a definite limiting value, indicates that, by absorption of quanta of sufficient magnitude, activated molecules are produced which have a longer life than the period of a molecular vibration (10⁻¹⁴-10⁻¹³ sec.), although shorter than that of a molecular rotation (10⁻¹²-10⁻¹¹ sec.). From these observations it appears that molecules can exist, at least for a short time, when they have absorbed more than enough energy to decompose them completely.

On the other hand, experimental evidence shows decisively that photochemical change can only be brought about when light of less than a certain wave-length is employed for illumination. There is thus a definite photochemical threshold which defines the minimum energy of the absorbed quanta of light; and it is only when the quanta exceed this value that the molecule becomes unstable.

(d) Activation and Deactivation. The non-polar activation of a molecule may involve either a mere increase of its energy-content with formation of an excited molecule, or a complete dissociation into free atoms. In the former case the energy may be stored up as kinetic energy or as energy of rotation and vibration, or (for very brief periods) as energy of electronic excitation. Electronic energy is usually lost at once by radiation; and in some circumstances energy of vibration of the nuclei and energy of rotation of the molecule may be lost by radiation at the same time as a change of electronic energy. In general, however, the energy of activation of a molecule is lost by collision with another molecule, which is in a condition to

^{1&}quot; Structure des Molecules", 1925.

accept energy from it. Thus nitrogen is specially efficient in the deactivation of carbon monoxide, perhaps because of a similarity in the magnitude of the quanta of rotational and vibrational energy which the two molecules can absorb.

A modern picture of activation by dissociation into free atoms ¹ is to suppose that, at the moment of absorption by the molecule of quantized energy (corresponding to an absorption band in the molecular spectrum), the orbits of a pair of covalency electrons are suddenly modified so that each electron of the duplet becomes planetary to one atom only. Centrifugal force then separates the two nuclei, and any extra energy remains as kinetic energy in the separated atoms until a molecular collision takes place.

In general, a "three-body collision" with another molecule is necessary before two atoms can recombine, since otherwise the energy of recombination of the dissociated atoms cannot be dissipated. In those cases in which a molecule is able to absorb quanta of light which are much greater than the energy required to dissociate the molecule into neutral atoms, it has been found that the surplus energy corresponds to the energy of excitation of one of the atoms to a higher stationary state. These excited atoms may radiate energy in the form of light of longer wave-length than that originally absorbed, thus giving rise to fluorescence; but if the excited state is metastable, an after-glow may be produced. Thus the glowing gas known as "active nitrogen" contains:

- (i) Neutral atoms of nitrogen, N.
- (ii) Excited atoms of nitrogen, N* and N**, with energies of excitation of 53,000 and 83,000 cal. per gramatom.
- (iii) Excited molecules, N₂*, with an energy of excitation of 195,000 cal. per gram-molecule.

The after-glow is attributed primarily to the recombination of excited atoms of nitrogen with neutral atoms of nitrogen,

¹ Cathala, J. Chim. Physique, 1928, 25, 185.

in presence of neutral molecules of nitrogen to dissipate the surplus energy 1:

$$N^* + N + N_s = 2N_s + after-glow.$$

(e) Negative Catalysis of Chain-reactions. The concentration of active atoms in a photochemical change, as calculated from the absorption coefficient of the light, is only of the order of one molecule in 108, but this is quite sufficient for reaction to occur at a noticeable speed. Any substance which can eliminate these atoms (or deactivate the activated molecules in an energy-chain) will terminate the chains and thus play the part of a negative catalyst. Thus foreign substances, such as oxygen, which may be present as impurities or added deliberately, may act as inhibitors to photochemical reactions by terminating chain processes. These same substances are also potent in quenching fluorescence.²

The surfaces of the reaction-vessels, and the impurities condensed on them, may also play a similar part, since it has been found that even the order of a chemical change may be altered by careful drying, or by pre-heating in vacuo, of the reaction-vessel. Conversely, it has been suggested that reaction-chains frequently commence at the surface of the vessel, where other factors than mere absorption of light have to be taken into account, and then spread through the gaseous phase, probably to terminate eventually at a surface once more. Chain-reactions are, therefore, not necessarily due to the initial production of neutral atoms.

In fact, it has been shown that α -rays and electrons can initiate reactions of a very similar nature to photochemical reactions, but by an entirely different mechanism.

Energy chains, in which "hot molecules" play the part of neutral atoms, and hand on a surplus of energy from

¹ E. J. B. Willey, J. Chem. Sqc., 1927, 2831; 1928, 1620; 1930, 336, 1146.

² F. Perrin, Ann. de Physique, 1929 (x), 12, 169.

⁸ Gwyn Williams, J. Chem. Soc., 1932, 1747, 1758.

⁴ S. C. Lind, "The Chemical Effects of α-particles and Electrons" American Chemical Society Monograph Series, No. 2.

one molecule to another without being decomposed thereby, have been proposed by Christiansen,¹ and have often been invoked when no obvious reaction-chain could be devised.

An excellent example of negative catalysis of a chain-reaction is found in atmospheric "autoxidation", e.g. of liquid benzaldehyde or of aqueous sodium sulphite. This can be initiated by the action of light, giving rise to a chain-reaction in which the absorption of each quantum of light is followed by the oxidation of 10,000 molecules of the aldehyde. The oxidation can be inhibited by a large range of negative catalysts, which appear to act by destroying a peroxide by which the chain is propagated. This deduction is confirmed by the fact that when alcohol is used up for each chain that is initiated in the system.

The action of "anti-knocks" in checking detonation or "knocking" in an internal combustion engine may be explained on similar lines.5 The "knocking" has been attributed to the formation of a peroxide as an intermediate product of the oxidation of the gaseous fuel. The propagation of a chain-reaction, in which the peroxide plays an essential part, is responsible for the detonation; and substances, such as lead tetraethyl, Pb(C₂H₅)₄, or iron carbonyl, Fe(CO), which act as "anti-knocks", are believed to do so because they destroy the peroxide and thus break the chains before the reaction has become explosive in character. Negative catalysts, such as these, are obviously most effective when used to control the "branched-chain" reactions described in the following paragraphs, where the margin between combustion and detonation is particularly narrow.

¹ Trans. Faraday Soc., 1928, 24, 596, 714, cf. Bäckström, J.A.C.S., 1927, 49, 714; Trans. Faraday Soc., 1928, 24, 601.

² See Moureu and Dufraisse, Compt. rend., 1928, 186, 196, and many other papers by the same authors.

⁸ Bäckström, J.A.C.S., 1927, **49**, 1460.

⁴ Bäckström, Trans. Faraday Soc., 1928, 24, 601.

⁵ Egerton and Gates, "Theories of Anti-knock Action", J. Inst. Petroleum Techn., 1927, 13, 281.

(f) Branched Chains. Usually each activated molecule or atom taking part in a chain-reaction reproduces only one similar activated molecule or atom. In other words, if each molecule activates α -molecules, the factor α is equal to or less than unity, since there is a definite risk $(\mathbf{r} - \alpha)$ that the process of activation may fail. Christiansen and Kramers ¹ have pointed out, however, that α may be greater than unity. "The reaction-velocity can then very easily become infinite, and the reaction becomes explosive."

The theory of branched chains was applied by Semenoff 3 to explain the existence of a minimum pressure below which oxygen will not unite with phosphorus, whilst above it oxidation is rapid and accompanied by luminescence. He suggested that a reaction-chain, started by an oxygen atom, might produce an activated molecule of oxide of phosphorus carrying so much energy that it would dissociate a molecule of oxygen, giving rise to two oxygen atoms in place of one. The initial formation of a single chain would then lead to a very rapid combustion, as soon as the branching of the chains had released them from the control usually exercised by the deactivating influence of chain-breaking collisions with the walls of the vessel or with other atoms or molecules. Similarly, Hinshelwood and Thompson 4 have postulated a mechanism of the homogeneous combustion of hydrogen and oxygen, depending on an initial formation of hydrogen peroxide, which (between certain narrow limits of concentration) develops explosively by means of a branched-chain reaction of the type:

$$\begin{array}{c} H_{3} + O_{3} = H_{3}O_{3} \\ H_{3}O_{3} + H_{3} = 2H_{3}O^{*} \\ H_{3}O^{*} + O_{3} = O_{3}^{*} + H_{3}O \\ O_{3}^{*} + H_{3} = H_{3}O_{3}. \end{array}$$

In this scheme each molecule of hydrogen peroxide gives rise to two activated molecules of steam, H₂O*, each of which gives rise to an activated molecule of oxygen, from which

¹ Zeit. phys. Chem., 1923, 104, 451.
² Ibid., p. 465.
³ Zeit. für Physik, 1927, 46, 109.
⁴ Proc. Roy. Soc., 1929, A122, 610.

a fresh molecule of hydrogen peroxide is produced. At low pressures, when the activating collisions must be relatively few, the chains are kept under control by deactivating collisions with the walls of the containing vessel; at high pressures, it is suggested, the molecules of hydrogen peroxide destroy one another. Explosion is therefore limited to a well-defined range of pressures, on either side of which the process is under complete control.

Ionic Reactions of Organic Compounds.

Totally different experimental conditions to those required for the gaseous reactions discussed above prevail in regard to chemical reactions which take place in solution at low temperatures and in the dark, since many of these are obviously ionic in character, even when they are concerned with organic compounds which are not themselves electrolytes. This statement can be applied without hesitation, for instance, to the action of potassium hydroxide on bromine in dilute aqueous solution (compare p. 167), to the saponification of an ester by a caustic alkali, or to the conversion of a sulphonic ester into an acetic ester by the action of potassium acetate.

$$Br - Br + 2 \overset{+}{K} \overset{-}{O}H = \overset{+}{K} \overset{-}{Br} + \overset{+}{K} \overset{-}{O}Br + H_{\bullet}O$$

$$CH_{\bullet} \cdot CO \cdot OC_{\bullet}H_{\bullet} + \overset{+}{K} \overset{-}{O}H = CH_{\bullet} \cdot CO \cdot \overset{-}{O} \overset{+}{K} + C_{\bullet}H_{\bullet} - OH$$

 $C_8H_{17} \cdot O \cdot SO_4 \cdot C_7H_7 + K \cdot \overline{O} \cdot CO \cdot CH_4 = C_8H_{17} \cdot O \cdot CO \cdot CH_4 + K \cdot \overline{O} \cdot SO_4 \cdot C_7H_7$ In each case it is clear (i) that it is the ions of the reagent that are active in bringing about the change, and (ii) that the products include an ion derived from the original covalent molecule. It is, therefore, reasonable to suppose (i) that at some stage of the process the neutral molecule of the reactant is split into two oppositely-charged ions by the rupture of a covalent bond, (ii) that one of these ions combines with an ion of the reagent to form a neutral molecule, whilst (iii) the other ion persists in the form of a salt in association with the second (metallic) ion of the reagent. In actions such as these the ionic character of the mechanism appears to be established beyond dispute, although it is not necessary to suppose that the ions formed by the rupture of the covalent molecule have an independent existence like those of the reagent, since the action may depend on the formation of an intermediate complex or aggregate of ions, rather than on a process of complete ionic dissociation. The essential point of the ionic conception of chemical changes is that the bond which unites the two radicals of the covalent molecule must be supposed to break unsymmetrically to form a pair of oppositely-charged ions, rather than symmetrically to form a pair of neutral radicals.

The polar mechanism can obviously be applied to all reactions which are brought about by *ionized reagents in solution*, but it is also a possible mode of reaction of any substance, such as a weak acid or a phenol, which is itself capable of dissociating into ions.

The actual formation of ions from compounds of this type is established most conclusively when they yield conducting solutions, e.g. in water or in some other ionizing solvent; and, in the case of weak electrolytes, the degree of dissociation into ions can be deduced from the conductivity of the solution by the method of Arrhenius, as the ratio $(\alpha = \Lambda_v/\Lambda_\infty)$ of the equivalent conductivity at a given dilution (e.g. v equivalents per c.c.) to that at infinite dilution.

It is not necessary, however, for a substance to yield an appreciable proportion of stable ions on dissolution in order that it may undergo ionic reactions, since many substances which are not electrolytes nevertheless give rise to ions under the influence of a reagent, as in the case of the alcohols, which yield metallic derivatives, and are therefore regarded as capable of ionic reaction.

$$2C_2H_5$$
—OH + 2Na \rightarrow 2C₂H₅O Na + H₂

On the other hand, the absence of electrolytic conductivity does not prove the absence of ionization, but only the absence of dissociation into *free* ions, since ion-pairs can

only be electrolysed when dissociated, e.g. by the intrusion of a watery atmosphere round each ion.

$$\stackrel{+}{\text{Na}}\stackrel{-}{\text{Cl}} + \text{aq.} \rightleftharpoons \stackrel{+}{\text{Na}} \text{aq.} + \stackrel{-}{\text{Cl}} \text{aq.}$$

Ions may also be produced during the course of a molecular rearrangement of an organic compound, even though the ionized fragments may be incapable of independent existence (see Chapter XIV).¹

The behaviour of compounds like the halogens, which can react either in ionic or in non-ionic fashion, will depend upon whether the energy of activation, under given experimental conditions, is less for the ionic or for the non-ionic change. When the difference is not too great, either or both processes may occur. Thus, in the chlorination of toluene,2 it is possible that nuclear substitution (which is promoted by the influence of a polar catalyst) may be an ionic reaction, whilst the chlorination of the side-chain (which is promoted by the influence of sunlight) may be a non-ionic reaction. Ionic reactions differ from non-ionic reactions, however, in that the influence of the environment is dominant in determining both the energy required for the ionic dissociation of the molecule and the probability that a sufficient amount of energy will be available at the moment when the molecule of the reactant collides with the ions of the reagent.

Classification of Chemical Reactions.

Professor Lapworth has given 3 the following classification of dissociations in order of decreasing energy-requirements.

(a) Purely ionic.

$$H - Cl \rightleftharpoons H + Cl$$
 $Cl - Cl \rightleftharpoons Cl + Cl$
 $CH_3 - I \rightleftharpoons CH_3 + I$

¹ Compare Whitmore, J.A.C.S., 1932, 54, 3274; 1933, 55, 4153.

² Cf. Book and Eggert, Zeit. Elektrochem., 1923, 29, 521; Mason, Smale, Thompson and Wheeler, J. Chem. Soc., 1931, 3150.

³ Private communication.

This type of dissociation requires so much energy that it is almost unknown in practice.

(b) Neutral.

$$H$$
— $Cl \rightleftharpoons H + Cl$
 Cl — $Cl \rightleftharpoons Cl + Cl$

This is the normal method of dissociation in gases.

(c) Complex ionic.

H—Cl +
$$n(H_2O) \rightleftharpoons \overset{+}{H} (H_2O)_y + \overset{-}{Cl} (H_2O)_x$$

Cl—Cl + $n(H_2O) \rightleftharpoons HOCl + \overset{+}{H} (H_2O)_y + \overset{-}{Cl} (H_2O)_z$
or in general

$$X-Y + nA + mB = \overset{+}{X}A_n + \overset{-}{Y}B_m$$

This is the common method of dissociation in solution, where the association of the solvent with the charged radical is an important factor in determining the stability of the complex solvated ions.

Induction of Ionic Reaction.

J. J. Thomson in 1914 was probably the first to indicate clearly how ionic reactions could be promoted between molecules.

"If the atom is a mixture of negatively electrified corpuscles [i.e. electrons] and positive electricity, it will produce in its neighbourhood a field of electrical force. . . . These forces will be exerted by the atom not merely on the atoms which are associated with it in the molecule of a chemical compound, but also on the atoms in other molecules, giving rise to forces between the molecules, and producing thereby the intrinsic pressure and surface tension of liquids, latent heat of evaporation, cohesion of solids and liquids, the rigidity of solids. and so on." 1

"Let us now consider what conditions would have an influence on the process of molecular ionization. . . . Anything which increases the 'capacity' of the molecule will make it easier for the atoms in it to acquire opposite electrical charges. . . . Suppose, for example, that A, B represent the atoms in a molecule which when by themselves have so small an electrical capacity that the work required to charge up

¹ Phil. Mag., 1914, [vi] 27, 757.

the one positively, the other negatively, is more than the energy available from their difference in the electrochemical series. Suppose that there are other molecules in the neighbourhood, and that the electric force round AB was sufficient to drag some of these molecules up to either A or B or to both of them, producing a more complex system. The proximity of these molecules to A and B would increase the capacity of the system, and would therefore facilitate the passage of a corpuscle from one atom to another, so that these atoms might get charged in presence of these molecules, whereas if they were alone they might remain uncharged." 1

"Let us consider the conditions necessary for the occurrence

of intramolecular ionization.

"(1) There must be a considerable difference in the electrochemical properties of the atoms A and B, to furnish the energy necessary for the transference of a corpuscle from one atom to the other; in other words, A and B must be well separated in the electrochemical series.

 \tilde{A} (2) There must be a strong field of electric force round AB so as to attract the molecules M_1 , M_2 . It is not necessary that molecules should attach themselves to both the atoms in AB, although the ionization will occur the more readily

if they do.

"(3) The molecules, M₁, M₂, must be strongly attracted by the field exerted by AB; thus, if they produce electrical effects analogous to those produced by electric doublets, the

moments of these doublets must be large." 8

"We have seen that union with another molecule promotes intramolecular ionization; thus it might happen that the atoms in AB were uncharged when AB was isolated, but when it had attracted another molecule a corpuscle would pass between A and B so that these atoms would get charged. In this way the properties of AB might be entirely changed by union with another molecule. We can, I think, go further than this. May there not be cases in which the attached molecule or molecules act on AB in much the same way as the molecules of a solvent would act upon it in solution, so that the atoms of A and B, in addition to getting charged through the proximity of the molecules attached to AB, might get separated by a kind of intramolecular dissociation, and then rearrange themselves so as to be in equilibrium with the atom of M, the molecule attached to AB?"

"To take the most famous case in connection with the formation of molecular compounds, that of sal-ammoniac:

¹ Phil. Mag., 1914, [vi] 27, pp. 769-70. ² Ibid., pp. 770-1.

we may regard this as produced by a molecule of ammonia, which as we have seen has an exceedingly large electrostatic moment, being attracted by the electric field round the molecule of hydrochloric acid. This is only the first stage: may not the HCl, under the action of the molecule of ammonia, not merely have the H atom positively, the Cl atom negatively electrified, but the two torn apart from each other? The four H atoms, and the N and Cl atoms might then regroup themselves so as to be in equilibrium." ¹

Thomson's conclusions may be summarized as follows:

- (a) Ionic dissociation of a molecule depends on augmenting its dipole moment.
- (b) This dipole moment (which may be small or non-existent in an isolated molecule) may be augmented by the proximity of molecules which already possess strong dipole moments, or which are already ionized.
- (c) This augmentation of the dipole moment decreases the energy of ionization, which may thus be diminished to a point at which the transfer of an electron from one atom to another takes place, giving rise to an ionized (instead of a merely dipolar) molecule.
- (d) The ions produced by this process of electrostatic induction may interact with the inducing molecule, which then plays the part of a reagent, instead of a mere catalyst; thus a dipolar molecule of ammonia may not only promote the intramolecular ionization of a molecule of hydrogen chloride, but may actually unite with the hydrogen ion thus produced.

Contrast Between Ionic and Non-Ionic Reactions.

The conclusions reached in the present chapter may be summarized as follows:

(a) Non-ionic reactions are characteristic of gases, the small dielectric constants of which make the formation of ions very difficult, but can also occur in non-ionizing solvents of low dielectric constant such as benzene or carbon tetrachloride. Since these reactions depend on the symmetrical rupture of a covalent bond, the energy cannot be

derived from an electrostatic field; but it may be obtained from radiation, or from intermolecular collisions giving rise to a small proportion of molecules of high energy-content in accordance with Maxwell's law of distribution of energy. The final rupture of the molecule then depends on the concentration, in the vibrational energy of a particular bond, of a sufficient portion of the total energy of translation and vibration which the molecule has derived, e.g. from collision with other molecules.

(b) Ionic reactions are characteristic of solutions, and especially of those with high dielectric constants, e.g. aqueous solutions. Unlike the non-ionic reactions, they are influenced to a predominant extent by the electrical environment of the molecule. They therefore take place in a polar environment, which may be produced by (i) a polar reagent, such as sodium acetate, caustic potash or sulphuric acid, (ii) a polar catalyst such as an acid or base in cases of homogeneous catalysis, (iii) a polar surface such as copper, alumina or glass, in cases of heterogeneous catalysis.

CHAPTER VI

UNSATURATION

Single, Double and Triple Bonds.

An unsaturated compound is defined as one in which the maximum valency is not exerted by all the component atoms, so that there is still a residue of combining-power left in the molecule. The most conspicuous illustrations of unsaturation are to be found amongst the "free radicals" described in Chapter VII, where the unsatisfied valency of one of the atoms (as represented by a single unshared and uncoupled electron) produces chemical activity of a very high order. A large number of compounds are known, however, in which a milder type of unsaturation is produced by the unused valencies of two contiguous atoms in compounds such as ethylene and acetylene. These compounds can be derived from ethane by the successive elimination of pairs of neutral hydrogen atoms; but the surplus valencies thus created are not regarded as free. since it is assumed that they have been used to saturate one another, and have thus lost the extreme reactivity of the surplus valencies of "free radicals". This type of unsaturation is therefore represented by means of double bonds and triple bonds, as follows:

In terms of the electronic theory of valency, the removal of a neutral hydrogen atom. consisting of a proton and one

electron, from a carbon atom to which i^+ is linked by two shared electrons, leaves an odd electron on the atom of carbon; but when two contiguous atoms of carbon are treated in this way, the odd electrons do not remain free, but unite to form an additional duplet, which is shared between the two carbon atoms, in addition to the shared duplet of the single bond of ethane, thus:

$$\begin{array}{cccc} H & H & H & H \\ \vdots & \vdots & \ddots & \vdots & \vdots \\ C & H & H & H \end{array}$$

The double and triple bonds of organic chemistry were therefore formulated by G. N. Lewis as the equivalent of two or three pairs of shared electrons.

Van't Hoff, when using a tetrahedral model of the carbon atom, represented the formation of single, double and triple bonds as due to the sharing of an apex, an edge and a face of two contiguous tetrahedra. A quadruple bond cannot be formulated by means of these models, and is therefore regarded as impossible, in strict agreement with modern theory and experiment (p. 44). These models are of great significance to the stereochemist, since it is found that a very stable form of geometrical isomerism may result from the cis- and trans-orientation of the radicals in unsaturated compounds such as fumaric and maleic acids. Since this isomerism does not occur in saturated compounds, such as the substitution-derivatives of ethane, X-CH₂·CH₂·X, the conclusion is drawn that, whilst free rotation can readily take place about the single bond which unites the two atoms of carbon in ethane, this rotation is inhibited when the two atoms of carbon are united by a double bond in ethylene. The conditions under which this inhibition may be overcome are discussed in a subsequent chapter (pp. 138 -141).

The conception that double bonds are formed by the sharing of a common edge in two contiguous tetrahedra was extended in 1891 by Hantzsch and Werner to include

the unsaturated compounds of tervalent nitrogen, where, however, one of the apices of the nitrogen-tetrahedron was supposed to be vacant. This hypothesis was introduced in order to explain the isomerism of the α and β oximes (p. 352), which appeared to be of the same geometrical character as that of fumaric and maleic acids, and has been established by the preparation of optically-active isomers, in cases where the geometrical isomerism is accompanied by molecular dissymmetry as well as by measurements of dipole moments (p. 79).

The tetrahedral configuration of many other elements which have a valency or a co-ordination-number of four has been established by similar methods; but the formation of double bonds is perhaps a peculiar privilege of the lighter elements, since Kipping, on dehydrating the silicols, invariably obtained products analogous with the ethers instead of with the olefines, and was unable to prepare any compound with a double bond between silicon and carbon:

$$\begin{array}{c} {}_{2}\text{R}\cdot\text{CH}_{2}\cdot\text{SiH}_{2}\cdot\text{OH} \longrightarrow (\text{R}\cdot\text{CH}_{2}\cdot\text{SiH}_{2})_{2}\text{O}} \\ \text{nct} & \text{R}\cdot\text{CH}_{2}\cdot\text{SiH}_{2}\cdot\text{OH} \longrightarrow \text{R}\cdot\text{CH}\text{=-SiH}_{2} \end{array}$$

In this connection two other cases may be noticed in which the double bonds of the conventional formulæ cannot now be maintained.

- (i) The strong paramagnetism of elementary oxygen proves that the odd electrons of the two atoms of oxygen have not united to form a duplet, so that the gas is not represented correctly by the conventional formula O...O.
- (ii) Compounds which contain semi-polar bonds are not represented correctly by formulæ such as O—SCl₂, O—PCl₃, etc., since only one pair of electrons is shared between oxygen and sulphur or phosphorus.

Reactivity of Double Bonds.

Double bonds have more than a steric significance, since compounds containing them have special reactions. Thus

¹ J. Chem. Soc., 1907, **91**, 209; 1912, **101**, 2106-66; 1921, **119**, 830.

whereas molecules containing only single linkages can yield derivatives only by a process of *substitution*, e.g. of chlorine for hydrogen, compounds with double or triple linkages can yield derivatives by a process of *addition*. Thus ethylene will interact with the halogens, halogen acids, hydrogen, sulphuric acid, aliphatic diazenes, ozone and other substances to form addition-compounds in which the two molecules are united without the formation of any other product.

These addition reactions, and not the double bonds of graphic formulæ, provide the chemical evidence of unsatura-They indicate that, as regards its chemical reactions, a double bond is not the equivalent of two single bonds, representing a binding of special strength between the two atoms (as it does in relation to its vibration-spectrum), but is actually a point of weakness in the molecule. This result can be interpreted very readily by applying van't Hoff's theory of directed valencies. Thus if a double bond is regarded as a two-membered ring, we should expect it to be even more strained and therefore more unstable than the three-membered ring in trimethylene and its derivatives. Support for this application of von Baever's strain theory is afforded by the high energy-content of unsaturated molecules, as evidenced by their negative heats of formation. The following table affords an illustration:

TABLE I

Heats of Formation of Hydrocarbons
(Cal. per gm. molecule.)

Ethane C ₂ H ₆	+ 25,500	Propane Propylene Trimethylene (cyclopropane)	C ₃ H ₆	+ 30,360
Ethylene C ₂ H ₄	- 6,480		C ₃ H ₆	- 1,920
Acetylene C ₂ H ₂	- 54,860		C ₃ H ₆	- 8,520

In this table, however, trimethylene is seen to be more endothermic than propylene.

Although "unsaturated" molecules all exhibit addition reactions at their multiple linkages, yet both the type and the mode of the addition process varies from molecule to molecule. Thus whereas the >C=C< linkage of the olefines interacts readily with halogens, halogen acids and ozone, the >C=O linkage is attacked by none of these agents, but combines with substances like ammonia, bisulphites and hydrogen cyanide to which the >C=C< linkage is quite inert. The N=N, C=N and N=O linkages are again different, the only reaction common to them all being reduction by addition of two atoms of hydrogen. On the other hand, compounds which contain semi-polar linkages are not unsaturated in the sense of the definition given above, since there is no duplication of covalencies and therefore no tendency to form addition-compounds. Thus the >SO group in the sulphoxides and sulphones is not unsaturated and does not form addition-compounds, since it has the structure $> \bar{S} - \bar{O}$, and does not contain the double bond which was assigned to it in the obsolete formula >S--O. In the same way, only one of the NO groups of nitro-compounds is unsaturated, since the other contains a semi-polar bond, $\stackrel{+}{>}N-O$, and is no more prone to additionreactions than is ammonium chloride.1

Mechanism of Addition-Reactions.

In a previous chapter it has been pointed out that whilst reactions in gases generally depend on the symmetrical non-polar fission of a covalent bond, giving rise to a pair of neutral radicals, the vast majority of reactions in solution depend on an unsymmetrical or polar fission of the bond, giving rise to a pair of reactive ions with opposite electrical charges and sharply differentiated combining properties. In a similar way, addition-reactions in unsaturated systems may sometimes be due to the symmetrical rupture of one duplet of a double bond, giving rise to a pair of "free radicals" linked by a single bond as indicated by the formula —CH₂—CH₂—; but in other cases the process appears to depend on the "ionization" of one of the

¹ See, however, pp. 73 and 396.

covalent linkages of the double bond by the transfer of both electrons of a shared duplet to the same atom, as in the formula $\overset{+}{C}H_2$ — $\overset{-}{C}H_2$.

(a) Ionic Reactions of Double Bonds. The ionic character of addition-reactions is particularly obvious in the case of carbonyl compounds. Thus the characteristic reactions of aldehydes and ketones may be attributed to a polar activation of the carbonyl group, >C=O, to the semi-polar form, $\stackrel{+}{C}$ —O. In this process the oxygen becomes a negative pole, which is capable of reacting with kations such as H^+ , or with reagents which readily yield reactive kations. The carbon atom, on the other hand, becomes a positive pole, and is therefore able to react only with anions such as CN^- , or with reagents which readily yield active anions. Since an oxygen atom can provide a fairly stable

seat for a negative charge, as in the anions OH, OC₆H₅, whereas a carbon atom with an ionic charge is nearly always unstable, one finds that the carbon atom of the carbonyl compounds is the primary centre of attack. The carbonyl group, therefore, reacts readily with reagents containing active anions, but not so readily with reagents in which the kation is active and the anion inert. The halogens, and their derivatives such as HCl, NOCl, etc., which yield relatively stable anions have therefore no tendency to form addition-compounds with the carbonyl group.

A conclusive experimental proof of the ionic character of a typical addition-reaction has been provided by Lapworth, who has shown that the familiar cyanhydrin condensation can be arrested indefinitely in mixtures of an aldehyde or ketone with prussic acid, but can be initiated immediately by the addition of a mere trace of alkali. This action is therefore not a direct addition of HCN, but

of the CN ion only, the metallic ion of the cyanide remaining free throughout the action. Although, therefore, the final result of the interaction can be represented as in (I), the

¹ J. Chem. Soc., 1903, 83, 995.

crucial stage is the condensation shown in equation (II):

(II)
$$CH_3 \cdot CH : O + HCN \rightarrow [CH_3 \cdot \overset{\leftarrow}{C}H - \overset{\leftarrow}{O} + \overset{\leftarrow}{H} \overset{\leftarrow}{C}N] \rightarrow CH_3 \cdot CH \overset{O}{\subset}N$$

(III) $CH_3 \cdot CH : O + \overset{\leftarrow}{C}N \rightarrow CH_3 \cdot CH \overset{\rightarrow}{\subset}N$
(III) $CH_3 \cdot CH \overset{\leftarrow}{\setminus} \overset{\leftarrow}{O} + \overset{\leftarrow}{H} \rightarrow CH_3 \cdot CH \overset{\rightarrow}{\setminus} \vec{C}N$

This condensation is reversible, and would soon be stopped by the competition for the cyanogen ions of the metallic ions of the salt. It is therefore necessary, in order to permit the action to proceed, to regulate the alkalinity of the solution by working in presence of a "buffer" compound such as potassium bicarbonate, which has an alkaline reaction, but can nevertheless supply hydrogen ions at a practically constant pH value and so generate the cyanhydrin itself as in equation (III).

(b) Non-Ionic Reactions of Double Bonds. The evidence for non-ionic addition-reactions is much less striking than that which has been cited above as a proof of the ionic character of the cyanhydrin reaction. It is, however, not difficult to recognize in some of the reactions of unsaturated compounds the characteristics which are cited in the following chapter as typical of non-ionic reactions. Thus, the addition of an atom of metallic sodium to the molecule of a ketone such as benzophenone or of an olefine such as styrene, without liberation of hydrogen (p. 126), can be explained most readily by the production of a free neutral radical, possessing an odd unpaired electron (cf. Chapter VII). This process implies that one of the two shared duplets of an ethylenic linkage can be broken symmetrically, giving rise to two "free radicals", linked together by a single bond as in the formula -CH₂·CH₂-. Again, a nonpolar activation of olefines may be postulated in order to account for their photochemical reactions. There are therefore valid reasons for thinking that double bonds may be

activated in a non-polar as well as in a polar fashion, and that some of their reactions may be of the non-ionic type.

Addition-reactions may therefore be formulated in two ways, as follows:

The second method of formulation is that due originally to Kekulé and to Wislicenus. This type of mechanism is certainly not universal amongst the reactions of unsaturated compounds, but on the contrary it cannot be excluded, or even regarded as abnormal, in view of the non-ionic character of many of the reactions of saturated compounds. Indeed, as in the parallel case of the chlorination of toluene (p. 149), it might easily happen that both mechanisms could be brought into operation when appropriate experimental conditions were provided, e.g. a polar catalyst on the one hand and a photochemical activation on the other hand.

Bivalent Carbon.

Unsaturation of a type different from that exhibited by the compounds dealt with in the following pages, is exhibited by a number of compounds, which have been formulated as containing bivalent carbon (p. 43). They are related closely to carbon monoxide, from which they may be derived as follows:

- (i) the *imides*, >C=NR, as found in the *iso*cyanides.
- (ii) the oxime, >C=NOR, as found in the fulminates.

Compounds of this type tend to polymerize; but their chief characteristic is a special type of unsaturation, whereby

two univalent radicals or one bivalent radical are added to the same atom of carbon, e.g.

(i)
$$R \cdot N = C$$
 + $O = R \cdot N = C = O$
 $R \cdot N = C$ + $S = R \cdot N = C = S$

(ii) $R \cdot N = C$ + $HCl = R \cdot N = C$
 $R \cdot N = C$ + $CH_3CO \cdot Cl = R \cdot N = C$

(iii) $R \cdot N = C$ + $H_2O = R \cdot N = C$

Giving $R \cdot NH_2$ and $O \cdot CH \cdot OH$

(iv) $R \cdot N = C$ + $R' \cdot NH_2 = R \cdot N = C$
 H
 NHR

Nef ¹ formulated these compounds as derivatives of bivalent carbon, containing two "latent valencies", as indicated in the preceding formulæ. This view was accepted by Thiele, who explained the moderate stability of the compounds by the mutual satisfying of a pair of latent valency bonds, $R \cdot N = C \subset C$.

The electronic structure of the *iso*cyanides may be represented in two different ways:

I. R:N::C: or R·N=C
II. R:N:::C: or R·
$$\dot{N}$$
= \ddot{C}

The first formula, corresponding to Nef's, shows two neutral atoms of nitrogen and carbon, but, whereas the nitrogen carries a complete octet, the carbon has only a sextet of valency electrons. The second formula, based on the isosterism of nitrogen and carbon monoxide, was suggested

¹ Annalen, 1892, 270, 267; 1895, 287, 265.

² Annalen, 1899, 306, 87; 308, 333.

first by Langmuir 1 and afterwards adopted by G. N. Lewis 2 and by Sugden. 8

Their chemical properties can be explained equally well by either formula. Thus in Formula I, the carbon, which possesses a lone pair of electrons in a sextet of valencyelectrons, can act both as donor and as acceptor and is obviously unsaturated, whereas the nitrogen with a complete octet can only act as donor. In Formula II, the quadricovalent nitrogen (which is positively-charged as in the ammonium ion) cannot act either as acceptor or as donor, since it carries a complete octet and has no lone pair of electrons. The carbon, on the other hand, carries a lone pair of electrons, which enables it to act as donor, e.g. to an oxygen atom or to a proton; but it can also act as an acceptor by releasing its hold on a pair of shared electrons, and allowing them to return to the exclusive control of the nitrogen, which thus acquires a lone pair and becomes tervalent and neutral instead of quadrivalent and positively-charged. The action of hydrogen chloride on an isocyanide can then be formulated as proceeding in stages as follows:

$$R: \overset{\uparrow}{N}: :: \overset{\downarrow}{C}: + \overset{\downarrow}{H} \longrightarrow R: \overset{\uparrow}{N}: :: C: H$$

$$R: \overset{\downarrow}{N}: :: C: H \rightleftharpoons R: \overset{\downarrow}{N}: : \overset{\downarrow}{C}: H$$

$$R: \overset{\downarrow}{N}: : \overset{\downarrow}{C}: H + : \overset{\downarrow}{C}: \longrightarrow R: \overset{\downarrow}{N}: : \overset{\downarrow}{C}: \overset{\downarrow}{C}: H$$

Although the chemical evidence is very ambiguous, the correctness of Langmuir's electronic formula for the *iso*-cyanides has been established both by Lindemann ⁴ and by Sidgwick and his colleagues ⁵ by measurements of parachors and of dipole moments.

¹ J.A.C.S., 1919, **41**, 1543. ² "Valence", 1923, p. 127.

^{3&}quot; The Parachor and Valency", 1930, p. 171.

⁴ Lindemann, Ber., 1930, 63, 1650.
⁵ J. Chem. Soc., 1930, 1876; Chemical Reviews, 1931, 9, 77; "The Covalent Link in Chemistry", Baker Lectures, Cornell University, 1933, pp. 187-90.

TABLE II

Parachors of isoCyanides, RNC

R. Methyl		[P] obs. I 22·I	[P] calc. for R—N=C. 96·6	[P] calc. + - for R·N≡C II8·4
Ethyl		164	135.5	157.3
Phenyl		255.2	230.5	252.3
o-Tolyl		292.9	269.5	291.3
p-Tolyl		295.5	269.5	291.3
o-Anisyl		314.1	289.5	311.3
p-Anisyl	•	314.5	289.5	311.3

As the above table indicates, the parachors of the isocyanides correspond much more closely to the formulæ of Langmuir than to those of Nef, and the evidence from dipole moments is still more convincing. Thus, in the cyanides the dipole, of moment 3.84×10^{-18} e.s.u., is directed so that the nitrogen is the negative pole. In the isocyanides, on the other hand, it is directed so that the nitrogen forms the positive pole, and the magnitude of the moment is considerable $(3.6 \times 10^{-18} \text{ e.s.u.})$. The small moment (0.12×10^{-18}) of carbon monoxide can be accounted for by the neutralization of the normal moment, 2.7×10^{-18} e.s.u. of the ketonic group > 0

by the electric charges shown in the formula O=C.

It may be noticed at this stage that, when expressed by means of electronic formulæ, the cyanides and iso-cyanides

means of electronic formulæ, the cyanides and iso-cyanigive rise to an identical ion, as set out below:

R:C:::N:
$$\leftarrow$$
 H:C:::N: \rightleftharpoons :C:::N: \rightleftharpoons :C:::N:R

Alkyl Hydrogen Common Hydrogen cyanide. cyanide. cyanide. cyanide.

This explains why prussic acid, which is perhaps an equilibrium-mixture consisting almost exclusively of hydrogen cyanide (p. 317), gives two series of alkyl-derivatives, but only one series of salts.

¹ Compare Lowry, J. Chem. Soc., 1923, 123, 828.

110 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

Acetylenic Compounds.

Nef 1 and Lawrie 2 have suggested that acetylenic compounds may exist in two forms as follows:

$$H \cdot C \equiv C \cdot H \rightleftharpoons H$$
 $C = C$

This point of view is supported by the oxidation of acetylenes to ketenes or to their immediate products of decomposition, and also by certain anomalous additive reactions ²:

e.g.
$$C_2Br_2 + HI \longrightarrow Br_2C \longrightarrow CHBr_2 \longrightarrow CHBr_2 \longrightarrow COOH$$

The complete symmetry of the molecule of acetylene has been established in a most conclusive way by spectroscopic evidence, and the unsymmetrical structure can therefore only occur (if at all) as a labile isomer of high energy content ³ (p. 49).

On the other hand, De Laslo, from a study of the diffraction of electrons by the vapour, has shown that dichloroacetylene, C₂Cl₂, possesses the acetylidene structure, Cl

have linear structures I—C=C—I. The latter compounds are stable, but dichloroacetylene invariably explodes when brought together with air. The theory of Nef has thus received a physical justification.

Azides and Aliphatic Diazo Compounds.

Angeli,⁵ and later Thiele ⁶ and Staudinger,⁷ have pointed out that azides and the aliphatic diazo compounds resemble

¹ Annalen, 1897, 298, 332: 1899, 308, 325.

² Amer. Chem. J., 1906, 36, 489.

⁸ (Mrs.) E. H. Ingold, J. Chem. Soc., 1924, 125, 1528.

⁴ Trans. Faraday Soc., 1934, 30, 825.

⁵ Gazzetta, 1894, 24, II, 46. ⁶ Ber., 1911, 44, 2197.

⁷ Helv. Chim. Acta., 1922, 5, 87.

the isocyanides and the fulminates in being able to add on two univalent radicals to one atom.

e.g.
$$CH_2N_2 + PhMgBr \rightarrow CH_2N \longrightarrow N$$

MgBr

They have, therefore, favoured a chain formula for these substances in preference to the original cyclic formula (A) of Curtius ¹ which accounts satisfactorily for the methods of preparation of these substances, but would indicate that addition should follow the rupture of a bond and ought to occur at two points.

In all, three formulæ are possible, if one maintains the octet rule:

Evidence to discriminate between them has, however, been most inconclusive.

In 1921 ² Levene and Mileska reported that an optically active diazo-succinic ester could be prepared from *l*-aspartic ester, and their observation was supported by W. A. Noyes ³ and others, ⁴ who prepared a series of aliphatic diazo compounds, all of very low optical activity. Asymmetry of the diazo compounds cannot be explained by the symmetrical cyclic formula (A), but is possible on formula (C), which, as Kendall and Noyes ⁵ pointed out, contains a semi-polar bond structure analogous to that found in the sulphoxides.

Compare
$$R \subset R' \longrightarrow R' \cong N \cong N$$
 with $R \hookrightarrow C$

¹ J. pr. Chem., 1891, (2), 44, 554.

² J. Biol. Chem., 1921, 45, 592. ³ J.A.C.S., 1922, 44, 1798.

⁴ Lindemann, Wolter and Groger, Ber., 1930, 63, 2404.

⁵ J.A.C.S., 1926, **48**, 2404.

112 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

The supposedly-active diazo compounds all contained a carbonyl group adjacent to the diazo group, and consequently cyclic formulæ of the betaine type, such as:

were also possible alternatives. However, the small optical activity ($\alpha_D = + \text{ i.5}$) which has been reported is usually attributed to the presence of an impurity.¹

Physical evidence does not enable one to discriminate conclusively in favour of any one formula. Langmuir sonsidered that the azides were isosteric with the cyanates, and should therefore be represented by similar electronic formulæ, but on the other hand Sidgwick in 1929 concluded that the volatility of the organic azides and of the aliphatic diazo compounds was indicative of the completely covalent ring structure (A), rather than that of either of the formulæ possessing semi-polar bonds.

Measurements of parachors did not give conclusive evidence, since there is only a difference of 2% in the alternative calculated values, in between which lie the experimental figures. The heats of combustion of phenyl azide and of ethyl azidoacetate, however, are not in accordance with a structure possessing a three-membered ring.

X-ray investigation of the crystal structure of metallic azides 6 has shown that the azide anion undoubtedly has a rectilinear configuration, and must therefore be $\bar{N} = \bar{N}$. In crystals of cyanuric triazide again a

¹ Weissberger, Ber., 1931, 64, 2896; 1932, 65, 265.

² J.A.C.S., 1919, **41**, 1543. ⁸ J. Chem. Soc., 1929, 1108.

⁴ Lindemann and H. Thiele, Ber., 61, 1928, 1529; Lindemann, Wolter and Groger, Ber., 1930, 63, 702.

⁸ Roth and Müller, Ber., 1929, 62, 1190; Sidgwick, Sutton and Thomas, J. Chem. Soc., 1933, 411; Sidgwick, Trans. Faraday Soc. 1934, 30, 801.

⁶ Hendricks and Pauling, J.A.C.S., 1925, 47, 2904.

linear arrangement of nitrogen atoms is found.¹ Further, Brockway and Pauling have investigated the structure of methyl azide vapour by the electron diffraction method, and conclude that the three nitrogen atoms are in line.² One can therefore exclude the cyclic formula (A).

The two semi-polar formulæ should each have large dipole moments, but oriented in different directions:

From experimental measurements it has been found that the dipole moment of aromatic azides, $Ar-N_3$, is about 1.5×10^{-18} e.s.u. with the negative end directed away from the aromatic ring, as indicated by the arrow in the formula above, whilst for diazomethane derivatives it has a magnitude of about 1.4×10^{-18} oriented in a similar sense.³ This result is incompatible with either formula (B) or formula (C), since a semi-polar bond should always have associated with it a moment of at least 4×10^{-18} e.s.u.

The only hypothesis that can reconcile all the known facts is to suppose that organic azides consist of molecules having structures (B) and (C) in tautomeric equilibrium, though the time taken in the conversion of one form into the other must be less than 10⁻⁸ second.⁴ L. Pauling,⁵ however, has suggested that when one could assign to a molecule two alternative structures which have nearly the same energy content and a similar spatial arrangement, then its actual state will be that of a hybrid structure intermediate between those of the two graphically depictable forms. Any such molecule is said to be in a state of resonance, and no simple graphical or electronic formula can represent adequately its molecular configuration, though its energy content can be estimated by the theories of wave-mechanics.

¹ Knaggs, Trans. Faraday Soc., 1934, 30, 826.

² Proc. Nat. Acad. Sci., 1933. 19, 860.

³ Sidgwick, Sutton and Thomas, J. Chem. Soc., 1933, 411.

Sidgwick, Trans. Faraday Soc., 1934, 30, 801.

⁵ J.A.C.S., 1932, **54**, 988. Compare Sidgwick, J. Chem. Soc.. 1936, 533.

114 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

It is now considered that other substances besides the azides and the aliphatic diazenes can exist in a state of resonance (see pp. 393-396). The interchange of electronic charge between molecular structures in resonance with one another will lead to the stabilization of a molecule by decreasing the total energy that need be associated with the molecular orbitals of its electrons.

It is necessary to point out here that a sharp distinction must be drawn between the conception of resonance as used by Pauling, and by Ingold, and the notion of "tautomerism" which is considered in detail in Chapter XIII.

¹ Chemical Reviews, 1934, 15, 241-65. Ingold uses the term, "mesomeric effect" to describe the electronic changes that must be associated with resonance.

CHAPTER VII

FREE RADICALS AND THEIR NON-IONIC REACTIONS

Simple and Compound Radicals.

The term radical has been used in chemistry since the days of Lavoisier 1 to denote the components from which chemical compounds are built up. Most of these radicals were elements like hydrogen or oxygen, but Gay-Lussac in 1815 discovered in prussic acid a compound radical, which he called cyanogen, 2 because it simulated the halogens; and Ampère in 1816 3 recognized in ammonium 4 a compound radical which behaved like a metal.

Compound radicals are defined as "groups of atoms which behave like the single atoms of a simple radical", in that they persist unchanged in a series of chemical reactions. Thus, the ammonium radical persists throughout the whole series of ammonium salts and differs from an alkali metal only in those reactions in which it is destroyed by oxidation, etc. In the same way, the cyanogen radical persists throughout the series of simple and complex metallic cyan-

¹ See, for instance, his "Traité Elémentaire de Chimie", tr. Kerr, 1797, chap. VIII. "Of the Radical Principle of Water" (p. 135), "Thus water, besides the oxygen, contains another element as its constituent base or radical" (p. 141).

³ "There is a great analogy between prussic acid and hydrochloric and hydriodic acids like them, it contains a radical which combines with potassium: only this radical is compound where chlorine and iodine are simple. To designate the radical of prussic acid, I have adopted the name cyanogen" (abbreviated from Ann. de Chimie, 1815, 95, 161-2).

⁸ Ann. Chim. Phys., 1816, 2, 16 (footnote).

⁴ The name was given by Davy in 1808 to the "metal" in ammonium amalgam, and the symbol **Am** was proposed by Berzelius in 1823.

ides and in a large range of organic compounds; but it differs from a halogen in that it can be decomposed by hydrolysis, oxidation or reduction.

As long ago as 1817, Berzelius concluded that "In inorganic nature all oxidized bodies contain a simple radical, while all organic substances are oxides of compound radicals." 1 Organic chemistry was therefore regarded as the chemistry of compound radicals. Thus, alcohol and ether were formulated by Dumas and Boullay as hydrates of olefiant gas (or ethylene); and the compound ethers, such as acetic ether and sulphovinic acid, were formulated as salts of olefiant gas, "which plays the part of a very powerful alkali, endowed with a capacity of saturation equal to that of ammonia ".2 A more remarkable example of the application of the theory of compound radicals to organic chemistry was provided, however, when Liebig and Wöhler 3 in 1832 concluded that the benzovl radical, C₇H₈O₈ was present as a hydride in oil of bitter almonds, as a hydroxide in benzoic acid, as a chloride in benzoyl chloride, and as a cyanide in benzoyl cyanide.

Preparation of Compound Radicals.

The quasi-metallic ammonium radical was prepared in the form of an amalgam in 1808 (i) by Berzelius and Pontin, who electrolysed aqueous ammonia with a mercury kathode, and (ii) by Davy, who acted on ammonia or an ammonium salt with an amalgam of potassium, sodium or barium. In the same way, the non-metallic cyanogen radical was separated in 1815 as an inflammable gas by Gay-Lussac, who prepared it by heating mercuric cyanide. Many of the early investigators therefore considered that, provided the requisite experimental conditions were discovered, organic radicals could be prepared in a free state, just like an alkali metal or a halogen. This conclusion was strength-

¹ Textbook, 1817, I, 544. ² Ann. Chim. Phys., 1828, 37, 52.

³ Liebig's Ann. Pharm., 1832, 3, 249-82.

⁴ Gilbert's Annalen, 1810, 6, 247. ⁵ Works, V, 122.

⁶ Ann. de Chimie, 1815, 95, 172-99.

ened when Bunsen, in 1842,¹ showed that the cacodyl compounds contain the cacodyl radical, AsC₂H₀, which he regarded as a true organic element, since it possessed all the properties of a metal and was capable of existing in the free state, although the "element" was not in fact a metal but a transparent liquid boiling at 170°. When, however, Cannizzaro established once and for all the method of deducing molecular weights from determinations of vapour density, by means of Avogadro's hypothesis, it was realized that groups such as methyl or methylene did not persist in the free state, like the monatomic elements sodium or mercury, but united into double molecules at their instant of formation, like the simple inorganic radicals, hydrogen and chlorine, when set free by electrolysis.

$$2CH_3 \rightarrow C_2H_6$$
; $2CH_2 \rightarrow CH_2:CH_2$

compare

$$HCI = H + CI$$
; $H + H = H_2$; $CI + CI = CI_2$.

The persistence of free organic radicals was also out of harmony with the doctrine of the quadrivalency of carbon, put forward by Kekulé in 1858, which soon became an accepted dogma of organic chemistry, since the doctrine of constant valency number, which was inherent in it, was obviously incompatible with the tervalence of carbon in the methyl radical or with its bivalency in methylene. The possibility of isolating the compound radicals of organic chemistry in an unpolymerized form was therefore generally abandoned during the period in which structural organic chemistry was advancing most rapidly.

Preparation and Properties of Free Organic Radicals.

The constancy of the valency number of carbon was challenged by Nef in 1890, on the basis of experimental evidence derived from investigations of fulminic acid, isocyanides, and other compounds containing "bivalent carbon" (p. 107). Nevertheless, great interest was aroused by the wholly-unexpected discovery, made by Gomberg in

¹ Liebig's Ann. Chem. Pharm., 1842, 42, 27.

1900,1 that hexaphenylethane tends to dissociate at ordinary temperatures into the "free radical" triphenyl-methyl, containing reactive tervalent carbon.

$$C_2Ph_6 \rightleftharpoons 2CPh_3$$

During the present century Gomberg's conclusions have been firmly substantiated and, in addition to many types of carbon compounds, free radicals have been prepared which contain unsaturated atoms of nitrogen, oxygen, sulphur, arsenic, tin and lead.2

Free radicals are to-day best described by paraphrasing the definition given by Wieland 3 in 1915 as follows: Free radicals are complexes of abnormal valency, which possess additive properties, but do not carry an electrical charge and are not free ions.

The real existence of neutral free radicals, analogous to free atoms of a halogen or of an alkali metal, was rendered less obvious by the fact that Gomberg's first product, hexaphenylethane (which has been the principal referencecompound in many discussions), is capable of dissociation either into neutral radicals or into ions. On the Continent. particularly in Germany, dissociation into free radicals has been postulated by investigators such as Schlenk and Wieland, who have substantiated this point of view by a study of the chemical reactions of these substances. In England and America, ionic dissociation has been stressed by Gomberg, Kraus and others (see Chapter VIII), who have made use of measurements of electrical conductivity to prove their case.

These two points of view are not incompatible, since in this respect hexaphenylethane merely resembles a halogen. the molecules of which are dissociated into neutral atoms under the influence of light, but are decomposed into positive and negative ions by an ionic reagent such as caustic potash.

Ber., 1900, 33, 3150; J.A.C.S., 1900, 22, 757.
 P. Walden, "Radicaux Libres", 3rd Solvay Report, 1928; W. Schlenk, "Radicaux Libres", 4th Solvay Report, 1931.

⁸ Ber., 1915, 48, 1098.

As Walden has pointed out,1 the reactions of free neutral organic radicals resemble those of free atoms either of a halogen or of an alkali metal. Thus the free methyl radical (prepared by heating lead tetra-methyl in a rapid stream of hydrogen or nitrogen under low pressure in a quartz tube 2) combines directly with lead, antimony, tellurium and zinc: and measurements of reaction-velocity have shown further that the free radical unites instantly with the element by mere collision, without requiring any energy of activation. These radicals therefore show an intense reactivity, which is quite unlike that of methane, ethylene or any other molecular product of decomposition of lead methyl. The free radical has, however, a mean half-life period of only 0.006 second and loses all its reactivity when condensed in liquid air. The ethyl radical,3 prepared by heating lead tetra-ethyl under similar conditions, behaves in a similar way. The instability and intense reactivity of these radicals correspond with the behaviour of a free atom of a halogen, where the equilibrium lies entirely on the side of the undissociated diatomic molecule or of the free atoms of hydrogen, with which so many experiments have been made in recent years.4 On the other hand, some of the more complex radicals of the triphenylmethyl series have only a small tendency to polymerize, the equilibrium being predominantly in favour of the free radical. These radicals then resemble the alkali metals, where the existence of diatomic molecules has only recently been established by spectroscopic observations.

Magnetic Properties of Free Radicals.

Free atoms of monatomic elements, such as hydrogen or sodium, differ from diatomic molecules like those of hydrogen

¹ Walden, 3rd Solvay Report, 1928.

^{*} Paneth and Hofeditz, Ber., 1929, 62, 1355.

⁸ Paneth and Lautsch, Ber., 1931, 64, 2702.

⁴ Bonhoeffer, Zeit. phys. Chem., 113, 119, 1924; 1926, 119, 385; Bonhoeffer and Harteck, "Grundlagen der Photochemie", Dresden. 1933; Harteck, Trans. Faraday Soc., 1934, 30, 134.

gas, in that they contain an odd number of electrons. The quantized "spin energy" of the odd electron gives to the atom a definite magnetic moment, which can be measured directly by the method of Gerlach and Stern, by passing the atom through an inhomogeneous magnetic field. In this way it has been shown that the hydrogen atom possesses one Bohr magneton, and that the alkali metals, nitric oxide, and "active" nitrogen gas also have magnetic moments. On the other hand, atoms or molecules which contain an even number of electrons are found to have no magnetic moment; according to current theories of atomic structure their electrons are usually associated in closed groupings containing equal numbers of electrons with spins of opposite signs.

In 1923 G. N. Lewis ³ deduced that all compounds whose molecules contained an odd number of electrons should be paramagnetic, and more recently Kuhn ⁴ has suggested that paramagnetism should be regarded as a physical property diagnostic of free radicals.

In accord with the modern quantum theory of magnetism,⁵ one can regard the molecular mass susceptibility, χ_M , of any substance which is not ferromagnetic as the sum of three distinct terms:

$$\chi_m = \chi \tilde{d} + \chi_\mu + \chi_r$$
.

 χ_d is the diamagnetic contribution, and is due to the disturbance of electronic orbits by an applied magnetic field. It is always negative, and ranges in magnitude from -1.88 for helium gas to a few hundred units for a complex organic molecule.

 χ_{μ} is the contribution due to the presence in a molecule of a permanent magnetic moment. Since the magnetic

¹ Ann. Physik, 1924, 74, 673; for a general account, see R. Fraser, "Molecular Rays", Cambridge Press, 1931.

² Jackson and Broadway, Proc. Roy. Soc., 1930, A 127, 678.

^{3&}quot; Valence and the Structure of Atoms and Molecules", 1923, p. 148.

⁴⁴th Solvay Report, 1931.

⁵ van Vleck, "The theory of Electric and Magnetic Susceptibilities", Oxford University Press, 1932.

moment of a molecule is determined by the resultant angular momentum of all the electrons, it is restricted, by quantum theory, to certain definite values. Thus, for a polyatomic molecule, χ_{μ} is determined almost entirely by the nett spin S, and is, to a close approximation, given by the equation:

$$\chi_{\mu} = 1.242 \times 10^{5} \times \frac{4S(S + 1)}{T}$$

In a free radical containing one unpaired electron $S = \frac{1}{2}$, and hence at 20° C. $\chi_{\mu} = +$ 1270. This is many times larger than χ_d and of opposite sign.

The third term χ_r is due to "exchange forces" or interactions between the electrons of adjacent atoms in the molecule, and is so small that it can usually be neglected.

Sugden 1 has shown experimentally that the measurement of magnetic susceptibility is a convenient test for the existence of a free radical, and the following table is taken from his results:

TABLE I

Magnetic Susceptibilities at 20° C.

Sı	ibstance	•				Xx	Xd	χ_{μ}
He						1 ⋅88	- 1·88	0
H_2O						13	(-13)	0
C.H.					_	56	(-56)	0
CuCl,					+	1,252	- 40	1,292
NO					+	1,465	– 10	1,475
ClO.					+	1,341	- 25	1,366
α-napl	ithyl-	diphe	nyl-					
met	hyl ·				+	570 (in 7%	solution i	n benzene)
Benzo	pheno	ne-po	tassiu	m				n in dioxan)
Phenyl-p-diphenylyl								
keto					+	1,080 (in 1	7% solution	on in dioxan)
O. (ga	.s) Î				+	3,310 to	+ 3,480	(calculated for
- 10	,					S = I	; + 3,390)	ì

Types of Free Radicals.

(a) The Pseudo-halogens. In addition to cyanogen gas, a number of substances are now known which may be classified as pseudo-halogens, e.g. oxycyanogen (CNO)₂,

¹ Trans. Faraday Soc., 1934, 30, 18.

thiocyanogen (CNS), tellurocyanogen (CNTe), all of which are liberated at the anode on electrolysis of their salts. These organic radicals resemble the halogens in that they combine with hydrogen to give strong acids, with metals to form salts (of which the mercurous silver and lead salts are usually insoluble in water) and with olefines to form addition-compounds. If dissociated in the same way as iodine, they would obviously yield free radicals; but as vet there is no decisive evidence to prove that this dissociation occurs. Nitrogen peroxide and perhaps nitric oxide (which resemble the free radicals in that they contain an odd number of electrons and possess marked additive properties), also show some resemblance to the pseudohalogens, as may be seen if we compare $N_2O_4 \rightleftharpoons 2NO_2$ with I, = 2I and KNO, with KI. Several analogues of these oxides are known amongst organic nitrogen-compounds (p. 125).

- (b) Pseudo-metals. The ammonium radical and its alkyl derivatives, such as N(C₂H₅), are classed as pseudometals. They can be prepared in the form of amalgams, or of blue solutions, by electrolysis of their salts in ether or liquid ammonia.1 The radicals of phosphonium, sulphonium and arsonium salts have also been isolated as amalgams. These pseudo-metals resemble the alkali metals in that they combine instantly with iodine or sulphur, and decompose water with liberation of hydrogen. Even the alkali metals, however, only behave as free radicals when in the state of vapour, since the magnetic properties of the atoms disappear in the solid metals. It is therefore probable that the pseudo-metals (like the pseudo-halogens) must be regarded as molecules, formed by the union of a compound radical with mercury or with ammonia, rather than as free pseudo-atomic radicals.
- (c) Triphenylmethyl and its Analogues. Triphenylmethyl was prepared by Gomberg 2 as an orange-yellow solution by abstracting chlorine by means of zinc from an

¹ Schlubach and Ballauf, Ber., 1921, 54, 2811.

² Ber., 1900, **33**, 3150; J.A.C.S., 1900, **22**, 757.

air-free solution of triphenylmethyl chloride in benzene. Triphenylmethyl can also be produced photochemically, since the system

$$Ph_3C-Cl + Ph_3C-H \xrightarrow{light} 2(Ph_3C\cdot) + HCl$$

gives rise to a reversible photochemical equilibrium.1

Its existence as a free radical was deduced from its great reactivity towards oxygen, iodine, sodium, and nitric oxide, all of which combine instantly with the product in a manner quite unexpected in a hydrocarbon. Thus, when a solution of triphenylmethyl is shaken with bromine or iodine it yields addition-products, namely, triphenylmethyl bromide or iodide, which give conducting solutions in liquid sulphur dioxide (p. 187). Triphenylmethyl also combines with metallic sodium to give a red solid, which yields conducting solutions in liquid ammonia (p. 192). It therefore bears a close resemblance both to the alkali metals and to the halogens, since it functions as a kation in the chloride and as an anion in the sodium derivative

$$CPh_3Cl \rightleftharpoons \overset{+}{C}Ph_3 + \overset{-}{Cl}$$
; $NaCPh_3 \rightleftharpoons \overset{+}{Na} + \overset{-}{C}Ph_3$.

In non-ionizing solvents, however, triphenylmethyl iodide dissociates to the extent of 20-40% into the neutral radical and molecular iodine ² just as molecular iodine dissociates into neutral atoms when heated.

$$_2\text{Ph}_3\text{C}-I \rightleftharpoons _2\text{Ph}_3\text{C} \cdot + I_2;$$

compare

$$I_2 \rightleftharpoons 2I$$
.

Triphenylmethyl combines directly with the free radical nitric oxide, to yield a stable product Ph₃C—N=O in which the two free valencies have neutralized each other. The characteristic reaction with atmospheric oxygen to form triphenylmethyl peroxide, Ph₃C—O—O—CPh₃, can also be classed as a combination of free neutral radicals, since

¹ Schlenk and Herzenstein, Ber., 1910, 43, 3545.

² Gomberg, Chemical Reviews, 1924, 1, 123; Gomberg and Nishida, J.A.C.S., 1923, 45, 1203.

oxygen is a paramagnetic substance which probably contains two unpaired electrons in each molecule (see p. 150). These addition reactions take place in solvents such as ether, petroleum and benzene, which are generally regarded as inert, because they do not promote reaction between polar compounds.

Determinations of molecular weight by measurement of freezing-point and boiling-point show that Gomberg's expected product, hexaphenylethane, is largely dissociated into its two component radicals, both in ionizing solvents such as liquid sulphur dioxide and in non-ionizing solvents such as benzene or naphthalene. The ionizing solvents give conducting solutions ¹

$$C_{2}Ph_{6} \rightleftharpoons \overset{+}{C}Ph_{3} + \overset{-}{C}Ph_{3};$$

but in non-ionizing solutions the neutral free radical is the only product of dissociation. The colours of the dissociation products in ionizing and non-ionizing solvents are therefore different.²

Amongst derivatives of triphenylmethyl, the degree of dissociation varies with the character of the substituents and increases with temperature. Moreover, some of the more complex compounds of this type, such as tridiphenylyl-methyl and pentaphenylethyl, CPh₃·CPh₃·CPh₄—, appear to exist wholly in the form of free radicals, and even form deeply-coloured crystals.

(d) Free Radicals containing Bivalent and Quadrivalent Nitrogen. Wieland in 1911 6 discovered that tetraphenylhydrazine dissociates to the extent of 10% in ben-

¹ Walden, 3rd Solvay Report, 1928.

² Gomberg and Sullivan, J.A.C.S., 1922, 44, 1811.

³ For discussion of this topic, see, *inter alia*, Marvell, J.A.C.S., 1930, **52**, 2976; Conant, J.A.C.S., 1925, **47**, 3068; 1929, **51**, 1925; 1933, **55**, 2098; Ziegler and Boye, Annalen, 1927, **458**, 248; Ziegler and Ewald, Annalen, 1929, **473**, 163; Ziegler, Trans. Faraday Soc., 1934, **30**, 10.

Schlenk, Annalen, 1910, 372, 1; Ber., 1910, 43, 1754.

⁵ Schlenk and Mark, Ber., 1922, 55, 2285.

⁶ Annalen, 1911, 381, 200; Wieland and Offenbacher, Ber., 1914, 47, 2113; K. H. Meyer and Gottlieb-Billroth, Ber., 1919, 52, 1476.

zene and 20% in nitrobenzene to give a coloured solution of great reactivity. Thus, with nitric oxide it gives diphenyl-nitrosamine, Ph₂N—N—O, with sodium the addition-product Ph₂NNa, and with triphenylmethyl the addition-compound Ph₃C—NPh₂. All these reactions indicate that the free radical, Ph₂N·, with an atom of bivalent nitrogen is capable of independent existence. This radical may be regarded as an organic analogue of nitric oxide, O—N·, in which the atom of oxygen has been replaced by two phenyl groups.

By the action of silver oxide on $\beta\beta$ -diphenylhydroxylamine, an atom of hydrogen is eliminated and the hydroxycompound is converted into diphenylnitric oxide,

$$Ph_2N \cdot OH \longrightarrow Ph_2NO.$$

This compound exists in solution exclusively in the monomolecular form. It can be regarded as an analogue of nitrogen peroxide, and resembles it in some of its reactions. Thus it liberates iodine from an acidified solution of potassium iodide, and is reduced by phenylhydrazine to diphenylhydroxylamine and by stronger reducing agents to diphenylamine

$$Ph_2NO \longrightarrow Ph_2N \cdot OH \longrightarrow Ph_2NH.$$

Dilute hydrochloric acid converts it by simultaneous oxidation and reduction into diphenylamine and quinone-anil, and a similar change sometimes occurs spontaneously in ethereal solution ¹

$$_{2}\text{Ph}_{2}\text{NO} \longrightarrow \text{Ph}_{2}\text{NH} + \text{Ph} \cdot \overset{+}{\text{N}} : \text{C}_{6}\text{H}_{4} : \text{O}.$$

A red product of similar type was prepared by Banfield and Kenyon² by the action of silver oxide on a condensation-product of β -phenylhydroxylamine with acetone

¹ Wieland and Roth, Ber., 1920, 53, 215.

² J. Chem. Soc., 1926, 1612.

Kenyon and Sugden ¹ have shown that, whilst the former compound is diamagnetic, like all ordinary organic compounds, the latter compound is paramagnetic and its molecule has a magnetic moment very close to that predicted for one unbalanced electron, namely $\mu = 1.75$ Bohr magnetons.

These red compounds have been formulated as containing quadrivalent nitrogen, e.g. Ph₂N:O; but this type of formula implies the existence of a shell of nine electrons round the atom of nitrogen, and it is therefore much more likely that the compounds contain an unsaturated atom of univalent oxygen, e.g. Ph₂N·O—.

(e) Free Radicals Containing Univalent Oxygen. Compounds which undoubtedly contain univalent oxygen can be prepared by the action of dry silver oxide on phenols, such as guaiacol, when reactive coloured solutions of the free radical are produced. A free radical of this type, discovered by Goldschmidt 2 in 1924, was dissociated to the extent of 69% at room temperature, and both the blue "free radical" and the colourless peroxide were isolated in the solid state

Cf) The Metallic Ketyls. Schlenk and Weickel in 1911 concluded that the sodium derivatives of the diaryl ketones, prepared by Beckmann and Paul in 1891, contained tervalent carbon, as in the formula — CPh₂·ONa. Similar products were then prepared by the action of metallic sodium and potassium on solutions of phenyldiphenyl ketone $C_0H_3\cdot CO\cdot C_0H_4\cdot C_0H_4$ and di-diphenyl ketone $(C_0H_3\cdot C_0H_4)\cdot CO\cdot C_0H_4\cdot C_0H_4$ and di-diphenyl ketone $(C_0H_3\cdot C_0H_4)\cdot CO\cdot C_0H_4\cdot C_0H_4)\cdot CO\cdot C_0H_4\cdot C_0H_4$

¹ J. Chem. Soc., 1932, 170.

⁸ Ber., 1911, 44, 1282.

² Annalen, 1924, 436, 202.

⁴ Annalen, 1891, 266, 1-28.

in ether or benzene. On exposure to air the sodium derivative of phenyldiphenylketone at once lost its green colour and was oxidized to sodium peroxide and the original ketone; with iodine it gave sodium iodide and the ketone. By the action of water, the metallic derivatives appear to give the hydrol —CPh₂·OH, which either polymerizes to form a pinacone, HO·CPh₂·CPh₂·OH, or gives rise by internal oxidation and reduction to a mixture of ketone and carbinol

$$_2$$
 —CPh₂OH \rightarrow Ph CHOH.

Subsequently 1 it was found that metallic sodium reacted with compounds containing the unsaturated groups, C=C, C=N, N=N, etc., attached to aromatic nuclei. An inorganic analogue, Na₂NO₂, has been obtained by Zintl and Kohn,² by treating sodium nitrite in liquid ammonia with metallic sodium.

Electronic Structure of Metallic Addition-Compounds. The alkali metal in the addition-compounds described above probably exists in the form of a kation (α) .

Ph
$$C \stackrel{-}{\longrightarrow} O$$
 Na; $C \stackrel{\beta}{\longrightarrow} O$ Na $C \stackrel{\beta}{\longrightarrow} O$ N

This is then attached by an electrovalent link to a negatively-charged atom (β) of oxygen, nitrogen or carbon, which is connected in turn with an atom having a neutral free valency (γ) . Thus the colour and great reactivity of the metallic ketyls recall the properties of organo-metallic compounds such as sodium triphenylmethyl, a compound which, like potassium pyrrole, is shown to be ionic by the electrical conductivity of its solutions in ether, pyridine or liquid ammonia (p. 192).

The free valency in the metallic ketyls and their analogues, appears on an atom which is not associated directly with

¹ Schlenk and Bergmann, Annalen, 1928, 463, 1.

² Ber., 1928, 61, 189.

the alkali metal; it has presumably been developed by the symmetrical fission of a covalent bond, followed by transfer of a single electron from sodium to oxygen, whereby both atoms are converted into ions.

$$CPh_2=O \longrightarrow \cdot CPh_2-O \cdot ;$$

$$\cdot \text{CPh}_2$$
—O $\cdot + \cdot \text{Na} \longrightarrow \cdot \text{CPh}_2$ —O Na

This method of fission of the double bond may be represented electronically as follows:

An alternative method of formulating the addition of an alkali metal to the double bond of an unsaturated hydrocarbon such as styrene has been suggested by R. Kuhn,¹ who supposes that a single electron can be transferred, resulting in the formation of a three-electron bond, which is capable of free rotation (II).

II.
$$K \cdot + \overset{\text{Ph}}{C} : \overset{\text{H}}{C} \longrightarrow K \cdot \overset{\text{Ph}}{C} : \overset{\text{H}}{C}$$

$$\overset{\text{H}}{H} \overset{\text{H}}{H} \overset{\text{H}}{H} \overset{\text{H}}{H}$$

$$\alpha \quad \beta \qquad \alpha \quad \beta$$

In this formula the alkali metal is held by a single-electron bond, similar to those which Sugden has postulated in compounds such as phosphorus pentachloride²; moreover, each atom of carbon has a complete octet of electrons, although one of them has a single unshared electron which, according to Kuhn, should behave like the odd electron in an alkali metal or in a halogen.

Either of the electronic formulæ I and II can be used to indicate the experimental fact that these metallic addition-compounds exhibit the reactions of free neutral radicals; thus:

(i) The metallic ketyls react readily with a further proportion of alkali metal to yield much more stable coloured solids, which in all their reactions resemble the metallic

^{1 4}th, Solvay Report, 1931.

alkyls and the Grignard reagents (compare Ph₃C·Na, HC=C·Na, Et₂Zn, CH₃·Mg·I).

(ii) The metallic alkylenes polymerize rapidly to yield double molecules which again resemble the metallic alkyls, whilst the metallic ketyls are easily transformed into pinacones.

(iii) The metallic ketyls oxidize rapidly to peroxides on exposure to air.

These are typical additive reactions of free neutral radicals, and can be expressed only by means of graphical formulæ containing free valency bonds, or by electronic formulæ containing an odd number of electrons.

Non-polar Fission of Ethylenic Linkages.

(a) The formation of the metallic ketyls, and of sodium derivatives of olefines such as styrene, can be explained most readily by the production of a free neutral radical, possessing an odd unpaired electron, as described above. This process implies that one of the two shared duplets of an ethylenic linkage can be broken symmetrically, giving rise to two "free radicals", linked together by a single bond as in the formula —CH₂—CH₂—.

A similar non-polar activation of olefines must be postulated in order to account for the action of light in promoting (i) addition-reactions, and (ii) geometrical isomerization (see below, pp. 138–142), since it is unlikely that photochemical activation of a symmetrical molecule would give rise to a pair of linked ions.

(b) The electronic structure of this system of linked radicals is shown in formula III, whilst that of the activated molecule postulated by Kuhn is shown in formula IV.

Formula III indicates clearly that the activated molecule

should possess two characteristic properties, namely, (i) two centres of reactivity and (ii) free rotation between the carbon atoms; but, since Kuhn has assigned similar attributes to formula IV, it is not easy to discriminate between the two formulæ. The fact that the absorption-spectrum of ethylene is characterized by alternating intensities 1 indicates that the molecules which absorb the light are completely symmetrical; but it obviously does not exclude the possibility that unsymmetrical activated molecules may be developed, e.g. in solutions containing an ionized electrolyte (p. 174) or a reagent with an odd electron.

(c) An activated molecule containing two linked radicals, which cannot separate freely like the free atoms of a dissociated halogen, would be expected to exist only momentarily, and to be converted almost immediately, by the recombination of the free valencies of the two radicals, into a stable molecule containing an ordinary double bond. Ingold and Marshall, however, have obtained definite evidence of the intramolecular production of free radicals by observing the changes of colour which take place on heating 9: 10-diphenylanthracene.

¹ Snow and Allsopp, Trans. Faraday Soc., 1934, 30, 93.

² J. Chem. Soc., 1926, 3080.

³ The central aromatic ring is shown with a para-linkage, which is apparently 3 A.U. in length, but this is not essential to the argument.

The colourless hydrocarbon forms coloured solutions in xylene; this colour becomes deeper on heating, but fades to its original depth on cooling. The addition of sodium or potassium to a solution in liquid ammonia gives a cherry-red solution, which deposits a brick-red precipitate containing two equivalents of metal. The addition-compounds with halogens also resemble those of triphenylmethyl in that they interact with water to form carbinols and with bases to form amines. Finally, oxidation and reduction proceed very easily at the 9:10 positions.

$$C_{\bullet}H_{\bullet} \xrightarrow{CPh} C_{\bullet}H_{\bullet} \longrightarrow C_{\bullet}H_{\bullet} \xrightarrow{CPh} C_{\bullet}H_{\bullet}$$

2:3:6:7-dibenzanthracene, which undergoes similar reactions, probably exists entirely in the free radical form, termed by Clar the "R"-form.

It is a deep-blue compound and its solutions do not darken further on heating.

The photochemical activation of anthracene itself, through which the polymer dianthracene can be obtained, is probably another example of the same type of dissociation. In these compounds a covalent bond can easily be broken on account of conditions of molecular structure

¹ Clar nd John, Ber., 1930, 63, 2967.

which, as in triphenylmethyl, favour the stability of free radicals.

Non-polar Addition-Reactions of Unsaturated Compounds.

The admission that an ethylenic linkage may be activated in a non-polar fashion implies that addition to the ethylenic linkage may also be non-polar in mechanism.

Thus the photochemical reduction of ethylene 1 and the halogenation of cinnamic acid and its derivatives (p. 138) definitely occur under non-polar conditions; but others, as for example chlorination and bromination of olefines in solution, or on polar surfaces, can be more easily interpreted by a polar mechanism (Chap. VIII).

Similarly, addition-reactions to "bivalent" carbon compounds such as carbon monoxide or *iso*-nitriles may be either polar or non-polar according to circumstances. The photochemical formation of phosgene from carbon monoxide and chlorine gases, for example, is certainly a chain reaction:

$$\begin{array}{ccc} \text{Cl}_2 & + h\nu & \longrightarrow 2\text{Cl} \\ \text{CO} & + \text{Cl} & \longrightarrow \text{COCl} \\ \text{COCl} & + \text{Cl}_2 & \longrightarrow \text{COCl}_2 & + \text{Cl} \end{array}$$

It is accompanied by side reactions, but the initial decomposition of the chlorine molecules into atoms is inferred from the connection which has been found experimentally between the measured rates of change with the concentration of the reactants, and the intensity of the illumination.² At high temperatures, for example, the reaction becomes reversible and the rate of change approaches more and more to a square-root relation with the concentration of the chlorine molecule, indicating that chlorine must initially split up into two active radicals which can both react at a measurable speed with the carbon monoxide.² On the other hand, a polar mechanism for the addition-reactions

¹ Taylor and Marshall, J. Phys. Chem., 1925, 29, 1140.

² Bodenstein, Zeit. phys. Chem., 1927, 130, 422; Bodenstein and Onoda, Zeit. phys. Chem., 1928, 131, 153; Bodenstein, Lenher and Wagner, Zeit. phys. Chem., 1929, B3, 459; Lenher and Rollefson, J.A.C.S., 1930, 52, 500.

of bivalent carbon compounds, which is consistent with their dipolar properties, has been advanced recently by Sidgwick and Hammick (Chap. VI), in order to explain the addition of ionized reagents to these unsaturated compounds.

Care must be taken therefore to decide upon the mechanism of any reaction only after a careful scrutiny of the actual experimental conditions in which it occurs. An attempt to illustrate the significance of these experimental conditions upon the mode of reaction is made in the following sections of this chapter, which deal with the more prominent examples of non-polar organic reactions.

Reactions of Sodium and Potassium Atoms.

(a) The action of neutral sodium atoms on various halogen compounds in the vapour phase has been studied by Polanyi, H. von Hartel and their colleagues.¹ Their experimental method depends on allowing sodium vapour, diluted with an inert gas, to stream from an aperture into a vessel containing a halogen compound at a temperature of about 300° C. and a pressure of 10⁻² to 10⁻³ mm. Under these conditions rapid reaction occurs, usually accompanied by chemi-luminescence in the form of a "highly attenuated flame". The chemical reactions which take place are illustrated by the following equations:

(i) Na + Cl₂ = NaCl + Cl; 2Cl = Cl₂;
(ii) Na + HCl = NaCl + H; 2H = H₂;
H + HCl = H₂ + Cl; 2Cl = Cl₂.
(iii) Na + CH₃Cl = NaCl + CH₃

$$CH_3 + CH_3 \rightarrow CH_3 - CH_3 \rightarrow CH_2:CH_2 + H_2;$$

 $2CH_3 + H_2 \rightarrow 2CH_4$

The initial action in equations (i) and (ii) requires no appreciable energy of activation, since every collision of a sodium

¹ Zeit. phys. Chem., 1928, **B1**, 21, 30, 61, 68; Zeit. phys. Chem., 1930, **B11**, 97; Zeit. phys. Chem., 1931, **B11**, 291; Zeit. phys. Chem., 1931, **B11**, 316; Trans. Faraday Soc., 1934, 30, 189.

² The sodium vapour was carried into the system in a stream of hydrogen; this would react with the free methyl to form methane, which was therefore the chief product.

atom with a halogen atom or with a halogen hydride is effective in producing chemical change. In these cases, therefore, the reaction velocity can be calculated from the mean speed of gaseous diffusion of the sodium vapour and checked by measuring the distance from the point of admission to which chemical action extends. This measurement can be made (i) by observing the dimensions of the attenuated flame, (ii) by measuring the area over which solid sodium halide is deposited, or (iii) by insolating the reaction mixture with a sodium flame and observing the area over which the sodium resonance spectrum can be detected.

With organic halides only a fraction of the collisions between sodium atoms and halide molecules results in chemical reaction. By assuming that the reaction constant **B** in the equation $k = \mathbf{B}.e^{-E/RT}$ is equal to the frequency of collision, it was possible to evaluate the combined kinetic energy E which must be attained by the two reactants before they can interact chemically. This deduction was verified in a few cases by calculating E from the rate of change of reaction velocity with temperature by means of the equation $d(\log k)/dT = E/RT^2$. The figures so obtained are given below:

TABLE II.

The Reaction between Sodium Vapour and Organic Halides

Halide.				Fraction of Effective Collisions.	Energy of Activation.
Methyl iodide			•	1:1.5	o cal.
Methyl bromide				1:25	3,200 cal.
Methyl chloride				1:5,000	8,800 cal.
Methyl fluoride				1:107	25,000 cal.
Ethyl iodide				1:5	1,700 cal.
Ethyl bromide				1:70	4,400 cal.
Ethyl chloride				1:1,000	7,300 cal.
n-Propyl iodide				I : 20	3,000 cal.
Iodobenzene				I:I	o cal.
Bromobenzene				I : 20	3,100 cal.
Chlorobenzene				1:1,000	7,200 cal.
Benzyl chloride		•		I : I	o cal.
Acetyl chloride				1:5	1,700 cal.
Benzoyl chloride				1:1	o cal.
Chloroform				1:6 ·	2,000 cal.
Carbon tetrachlo	ride			1:2	o cal.

From the above table it can be seen that the ease of decomposition of halides is I>Br>Cl>F. This is the same order as for their direct photochemical decomposition (p. 183). The nature of the organic radical has a much smaller influence on the energy required for the rupture of the carbon-halogen linkage.

The production of light appears to be a secondary phenomenon in those cases in which the energy of activation is large. It was attributed to the collision of sodium atoms with activated molecules, produced during the reaction and carrying sufficient kinetic energy to transmit to the alkali metal the energy required to excite radiation of one of its characteristic spectral lines.

(b) The high-temperature reactions discussed above are of interest in relation to the well-known reactions of Wurtz and of Fittig for the synthesis of hydrocarbons by the action of sodium on an alkyl iodide, and to Ullmann's reaction in which copper powder is used

$$2C_{2}H_{5}I + 2Na = C_{2}H_{5}-C_{2}H_{5} + 2NaI$$

 $2C_{4}H_{5}I + 2Cu = C_{4}H_{5}-C_{6}H_{5} + Cu_{2}I_{2}$

The analogy between these reactions suggests that the synthesis of the hydrocarbon may depend on the production of free alkyl radicals ¹; and this conclusion is confirmed by analogy with the action of sodium on triphenylmethyl chloride, where the relatively stable triphenylmethyl is actually produced as a neutral free radical. Alternatively, the synthesis may perhaps depend on the intermediate formation of a metallic alkyl:

$$C_2H_5I + 2Na = C_2H_5Na + NaI$$

 $C_2H_5Na + C_2H_5I = C_2H_5 - C_2H_5 + NaI$

The second stage of this process may then be compared with Williamson's synthesis of ethers:

$$C_2H_5O-Na + C_2H_5I = C_2H_5-O-C_2H_5 + NaI$$

¹Cf. Schlenk, 4th Solivay Report, 1931; Johnson and Adkins, J.A.C.S., 1931, 53, 1520; 1932, 54, 1943.

The first process proceeds by means of neutral alkyl radicals, although these are produced by eliminating a neutral iodine atom as an iodide ion by interaction with an atom of sodium. In the second process, the neutral alkyl radical interacts with a second atom of sodium, and is perhaps ionized thereby, since the final stage is almost certainly ionic:

$$C_2H_5Na + IC_2H_5 \longrightarrow C_2H_5 \cdot C_2H_5 + NaI$$

Displacement Reactions.

The inner mechanism of reactions involving metallic sodium, molecular silver, free iodine, or other atoms of elements is worth further consideration.

(a) Metals. When metallic sodium is used as a reagent, it always loses an electron and is converted into a stable sodium ion. The transfer of this electron to an acceptor molecule then yields a product containing an odd number of electrons. When, as is usual, a covalent bond is broken, the atom which received the odd electron becomes the seat of an ionic charge, whilst the atom adjacent to it gains the free neutral valency. Thus, in the formation of a metallic ketyl the oxygen receives an electron from the sodium, the covalent linkage between carbon and oxygen is broken symmetrically, and the carbon atom becomes the seat of the free valency.

$$Na + Ph_{2}C \longrightarrow Ph_{2}C \longrightarrow Ph_{2}C$$

Displacements of the same type, in which two products are formed instead of only one, are seen in the following familiar reactions:

$$CH_3$$
—I + Na = CH_3 · + Na I
 CH_3 ·OH + Na = CH_3 · ONa + H·

The solution of a metal in an acid with evolution of hydrogen also belongs to the same category, e.g.:

$$CH_3 \cdot CO \cdot OH + Na = CH_3 \cdot CO \cdot O \cdot Na + H.$$

The final products are, of course, the twinned radicals, H₂ and C₂H₄.

(b) Halogens. The metallic elements act as a source of the additional electrons that are required to break a covalent bond. Electronegative elements act in a converse manner by removing the electrons that become surplus when a covalent bond is formed. Thus the linking together of two hydrocarbon radicals by abstracting a metallic radical with a halogen is the true converse of the Wurtz reaction. As simple examples of this, the action of iodine on sodium thiosulphate, or on sodiomalonic ester, may be cited:

$$_{2}$$
Na $_{C}$ H(COOEt) $_{2}$ + $_{2}$ = (EtOCO) $_{2}$ CH—CH(COOEt) $_{2}$ + $_{2}$ Na $_{1}$

(c) **Electrolysis.** Displacement-reactions also occur in electrolysis, when the free atoms or neutral radicals produced by electron-discharge at the electrodes often take part immediately in secondary chemical changes. Thus the liberation of hydrogen at a kathode is often used to effect the reduction of organic compounds. The formation of symmetrical pinacones by the reduction of ketones may be cited as evidence that the hydrogen acts as a neutral radical in a non-ionic reaction:

$$\begin{array}{c} \text{CH}^3 \\ \text{C} \\$$

Kolbe's electrolytic synthesis, which has wide practical applications, may be cited as an analogous anodic change involving free neutral radicals. Thus the electrolysis of potassium acetate may be represented by the following scheme: 1

$$CH_3 \cdot CO \cdot OK \longrightarrow CH_3 CO \cdot O \longrightarrow CO_2 + CH_3 \longrightarrow CH_3 - CH_3$$

The initial stage in this action is obviously the loss of an electron when the acetate ion is discharged at the anode. The stable ion is thus converted into an unstable free radical with an odd number of electrons. This then decomposes

¹ Brown and Walker, Annalen, 1891, 261, 107; Walker, J..Chem. Soc., 1926, 3111; 1928, 2040.

into carbon dioxide and a methyl radical, which finally unites with a similar radical to form ethane.

Kolbe's electrolytic synthesis gives good yields only when very concentrated solutions are used. In dilute solution other actions predominate. These are in general oxidations, and have been traced to the action of hydrogen peroxide, formed by the discharge of hydroxyl anions: $2OH \rightarrow H_2O_8$.

Amongst other compounds methyl alcohol can be formed by electrolysis of an aqueous solution of an acetate.²

Interconversion of Geometrical Isomers.

The interconversion of camphoric and iso-camphoric acid in presence of concentrated hydrochloric acid, and the racemization of tartaric acid by strong soda, are obviously reactions of the ionic type, probably depending on a process of enolization (p. 309). Geometrical inversions can, however, be effected by an alternative non-ionic mechanism. with the help of catalysts (containing odd numbers of electrons) which appear to act as free radicals. Thus maleic ester can be converted into fumaric ester by the action of metallic potassium³; nitrous acid (containing NO) will convert oleic into elaidic acid; and, under the influence of light, halogens are resolved into free atoms which can initiate chain-reactions and are therefore very powerful catalysts for this type of change. The free atoms are true catalysts, since Latiers 4 has shown that there is no displacement of equilibrium in the photochemical inversion of the cis- and trans-di-iodoethylenes.

The halogens often yield addition-compounds as well as isomers; Wislicenus therefore suggested in 1887 a complex mechanism requiring (i) halogen addition, (ii) epimerization of the addition-product, and (iii) final re-elimination of the halogen. When it was shown that halogen-addition was only of the order of 5% or less, this theory was aban-

¹ Glasstone and Hickling, J. Chem. Soc., 1934, 1878; 1936, 820.

^{*} Hofer and Moest, Annalen, 1902, 323, 284.

³ Meerwein, Ber., 1926, 58, 1266.

⁴ Bull. Soc. Chim. Belg., 1922, 31, 73.

doned for the "affinity demand" hypothesis of Werner (1904); but this neglected to account for the rôle of the catalyst. To-day it is considered that halogen addition is a side reaction in a non-polar chain reaction process that is photochemically initiated.¹

The modern explanation of inversion of geometrical isomers, as developed by Berthoud and his colleagues from a whole series of researches,² is as follows:

- (i) Halogen activation.
- (ii) Atomic addition.



- (iii) Free rotation about the single bond.
- (iv) Elimination of the Br atom and deactivation.



Inversion (stage iii) is thus supposed to occur in a molecule containing a free radical and held together exclusively by single bonds. The process of "free rotation" probably occurs under the action of electrical forces exerted between dipolar molecules and would take place so as to reduce the total potential energy.⁸

This mechanism of transformation provides a quantitative interpretation of the results obtained by measuring the rates of inversion of such substances as maleic acid, stilbene and cinnamic acid. Thus, allocinnamic acid (the cis-isomer of cinnamic acid) can be converted to the ordinary trans-cinnamic acid, by exposing a solution in benzene to ultra-violet light, after adding a little bromine or iodine.

¹ Wachholtz, Zeit. phys. Chem., 1927, 125, 1.

² 3rd. Solvay Report, 1928, pp. 137-41; J. Chim. Phys., 1928, 25, 40; 1930, 27, 291; 1927, 24, 213.

³ Cf. Wolf, Trans. Faraday Soc., 1930, 26, 315; also chap. IV.

⁴ J. Chim. Phys., 1930, 27, 291.

Since transformation only occurs on exposure to light, the change of composition of the mixture can be followed by evaporating off the benzene in the dark, and comparing the melting-point of the residue with that of known mixtures of the geometrical isomers. It has been found (i) that the rate of transformation of allocinnamic acid under constant illumination is directly proportional to its concentration, and (ii) that the rate of transformation varies as the square root of the intensity of the light, indicating that each quantum of light absorbed must produce two active radicals, as in stage (i) of the scheme given above. When the light is completely absorbed, the rate of transformation is independent of the concentration of the catalyst (iodine), since the concentration of iodine atoms is limited by the amount of light that is absorbed. When more light energy is available than the solution can absorb, the rate of reaction is proportional to the square root of the concentration of iodine, since the number of molecules that are activated is proportional to the total concentration of iodine and each molecule that is activated supplies two catalysing atoms (stage ii). More complicated relationships are needed to explain the interconversions of α -phenyl- ω -cinnamo-nitrile, which is reversible 1; but in no case has any intervention of ionized substances been detected.

Trans-addition to the Ethylenic Linkage.

The formation of addition-compounds, as by-products in the interconversion of geometrical isomers, can be attributed to a side-reaction following stage (iii) of the scheme set out above, e.g.

¹ Berthoud and Nicolet, J. Chim. Phys., 1928, 25, 40.

In this side-reaction, the free radical (I), in which free rotation to the most stable configuration has occurred, combines with another atom of halogen, instead of losing its first halogen atom. The only requisite for trans- (in place of cis-) addition to any olefine is, therefore, that the two radicals (or ions) should be added independently, at a sufficient interval of time for the intermediate free radical (I) to assume its stable configuration. The mechanism already postulated for the geometrical interconversion therefore suffices to account for all cases of photochemical addition.¹

Addition of unsymmetrical substances, such as the halogen acids, or addition to highly unsymmetrical olefines, such as acrylic acid, may, however, proceed much more readily by a polar mechanism than by a non-polar process; and the halogenation of ethylene itself, when in polar solvents or in contact with polar surfaces, has actually been shown to take place by a different mechanism (Chap. VIII). It is indeed possible that addition of halogens proceeds as a non-polar action only under the influence of light. Should this be true, the difference between the two processes could be illustrated very well by the interaction of benzene with chlorine, for the action of sunlight yields the addition-product, C_eH_eCl_e, whilst in the presence of polar substances only substitution derivatives of benzene are produced.

Photo-polymerization.²

Polymerization of unsaturated compounds to double molecules can undoubtedly take place in the absence of any ions. Obvious examples are the formation of truxillic and truxinic acids from cinnamic acids and of dianthracene from anthracene. The former process always yields mixtures of isomers, which have no definite stereochemical relation to the geometrical configuration of the cinnamic or allo-cinnamic acid used to produce them, indicating that mixtures of free radicals have reacted. In confirmation of

¹ Berthoud, Helv., 1930, 13, 385.

For recent evidence see Trans. Faraday Soc., 1936, pp. 1-412.

this conclusion, Berthoud 1 exposed α -phenyl- ω -cinnamonitrile to the prolonged action of ultra-violet light, in presence of bromine, and obtained a polymer of the truxillic acid type, together with an intermediate compound containing two atoms of bromine

This intermediate compound had obviously been formed by condensation of two activated but non-ionized radicals of the type that has already been postulated in order to account for the processes of geometrical inversion and simultaneous addition of halogen.

Non-polar Dissociation.

As long ago as 1890 Nef ² suggested that many organic reactions might proceed through non-polar dissociations, with subsequent addition, after the manner of the dissociation of the permanent gases in the high-temperature syntheses of water, ammonia or hydrogen chloride. He stated that for each compound there existed a specific heat of activation which had to be reached before dissociation of the molecule into radicals could occur.

This hypothesis explained successfully the formation of diphenyl on passing benzene through a red-hot tube, and was applied to several other reactions in an analogous fashion.

Recently F. O. Rice 3 has confirmed this theory by showing that the thermal decomposition products of several

¹ Helv., 1930, 13, 385.

² Annalen, 1890, 258, 261; 1891, 266, 52 J.A.C.S., 1904, 26, 1549; cf. Henrich, "Theories of Organic Chemistry", chap. XIV (1922 translation by Johnson and Hahn).

⁸ J.A.C.S., 1931, 53, 1959; Chemical Reviews, 1932, 10, 135; Trans. Faraday Soc., 1934, 30, 152; J.A.C.S., 1932, 54, 3529, 3540. Compare also "The Aliphatic Free Radicals", by F. O. and K. K. Rice, Baltimore, 1935.

classes of organic compounds, such as hydrocarbons, ketones, aldehydes, amines and ethers are free radicals capable of reacting with mirrors of lead, zinc, antimony and mercury with the production of identifiable organometallic compounds. The usual product formed on "cracking" an organic compound at 450–750° C. is the radical methyl, for it appears that the more complex hydrocarbon radicals easily break down with the formation of an olefine, e.g.

$$CH_3-CH_3-CH_3$$
 \longrightarrow CH_3 $+$ CH_3

Rice has shown that these high-temperature decompositions are unimolecular reactions, for which the energies of activation lie between 50,000 and 67,000 calories, in good agreement with the theoretical estimate of the energy required to bring about a symmetrical fission of a carbon to carbon linkage.

Even unsymmetrical organic compounds can dissociate into free neutral radicals, just as hydrogen iodide can dissociate into its elements. Thus, methyl iodide and ethyl iodide have both been shown to decompose on exposure to light with separation of free iodine and hydrocarbon gases. Moreover, when liquid ethyl iodide is irradiated with light corresponding to an absorption band, Einstein's law of photochemical equivalence is very closely obeyed. The production of free alkyl radicals in this process has been established by obtaining a 35% yield of ethane from the products of photochemical decomposition of methyl iodide.

Dissociation of alkyl iodides into atomic iodine and an alkyl radical will also explain their photochemical oxidation to alcohols and aldehydes in presence of oxygen. A similar explanation, requiring non-ionic separation of hydrogen, has been advanced to explain the photochemical oxidation of ethyl alcohol by bromine. Thus the thermal

¹ Job and Emschwiller, Compt. rend., 1924, 179, 52; Franck, Trans. Faraday Soc., 1925 (discussion on photochemistry).

² Iredale, J. Phys. Chem., 1929, 33, 290.

⁸ Bates and Spence, J.A.C.S., 1931, 53, 1689.

⁴ Berthoud and Béraneck, J. Chim. Phys., 1928, 25, 40.

reaction in the dark, which is unimolecular with respect to the bromine molecule, can be expressed as follows:

$$C_2H_6O + Br_2 = C_2H_4O + 2HBr;$$

but it is more likely to be an ionic reaction, in which one active bromine ion (the Br⁺) is formed from each bromine molecule. On the other hand, the photochemical reaction proceeds at a rate which is proportional to the concentration of the alcohol, but to the square root of the concentration of bromine and of the intensity of the light. It is therefore formulated as a chain reaction:

(i) $Br_2 \rightarrow 2Br$ (ii) $C_2H_6O + Br \rightarrow C_2H_6O + HBr$ slow (iii) $C_2H_5O + Br_2 \rightarrow C_2H_4O + Br + HBr$ fast

in which the neutral bromine atom abstracts a neutral hydrogen atom, leaving a tree radical from which a second hydrogen atom can then be eliminated by molecular bromine.

The action of bromine on hydroxy-acids can also proceed by these two different processes, according as the interaction occurs in the dark or with absorption of light.¹

Decomposition of Azo-Compounds.

Decisive experimental evidence of the production of free neutral alkyl and aryl radicals, and of reactions brought about by them, has been obtained from the study of the decomposition of certain azo-compounds.

(a) Diazomethane decomposes explosively on heating to yield a complex mixture of gases.² It can be decomposed photochemically by the light of a mercury vapour lamp,³ and the nature of the decomposition products is only accounted for satisfactorily by supposing that the free radical methylene is produced:

$$h_{\nu} + H \longrightarrow C = N = N \longrightarrow CH_{s} + N_{s}$$

¹ Purkayastha and Ghosh, Zeit. phys. Chem., 1930, B7, 276.

² Steacie, J. Phys. Chem., 1931, 35, 1493.

⁸ Kirkbride and Norrish, J. Chem. Soc., 1933 119.

- (b) Azomethane, CH₃—N—N—CH₃, first prepared by Thiele,¹ decomposes on heating to give almost entirely nitrogen and ethane,² ethylene being produced only by a minor side-reaction. The decomposition is unimolecular and occurs in the gaseous phase, requiring no catalysts. Further, it can be photochemically activated. The gases obtained by heating azomethane at 450–550° will react with mirrors of antimony, and hence contain free methyl.³ Azopropane behaves similarly.⁴
- (c) Aromatic azo-compounds of the type of benzeneazo-triphenylmethyl, Ph-N=N-CPh, decompose gradually, with evolution of nitrogen gas, on heating either alone or in inert solvents, and from the residue Gomberg over thirty years ago isolated tetraphenylmethane, obviously produced by the union of the two hydrocarbon residues, together with some hexaphenylethane. Wieland and his collaborators, by carrying out the same reaction in presence of oxygen, have demonstrated the formation of from 10 to 30% of triphenylmethyl peroxide, which must have come from the free neutral radical, triphenylmethyl, produced in the change. The phenyl radical, instead of condensing to diphenyl, is reduced to benzene by other products of the reaction. Even in the absence of oxygen the formation of triphenylmethyl has been detected by its absorption spectrum; but the conclusions to be drawn from this decomposition are not decisive, since triphenylmethyl might be formed by a secondary decomposition of tetraphenylmethane or of hexaphenylethane.
- (d) More decisive evidence has been obtained ⁶ from the decomposition of acyl azo-compounds prepared by the oxidation of acyl hydrazides. Thus the hydrazide (II), when dissolved in chloroform, can be oxidized by potassium

¹ Ber., 1909, 42, 2575.

² Ramsperger, J.A.C.S., 1927, 49, 912, 1495.

³ Leermakers, J.A.C.S., 1933, 55, 3499; Rice and Evering, J.A.C.S., 1933, 55, 3898.

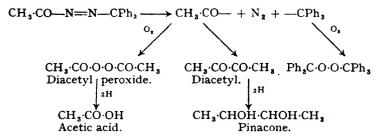
⁴ Ramsperger, J.A.C.S., 1928, 50, 714.

⁵ Ber., 1922, 55, 1816.
⁶ Annalen, 1927, 452, 1.

ferricyanide to the azo-compound (III), which can be isolated in orange-yellow needles. On warming the solution, however, nitrogen is set free and a cherry-red colour, like that of the free ketyls, is developed. On exposure to air, the red solution soon loses its colour and yields triphenylmethyl peroxide and the ketone IV.

$$\begin{array}{c} C_{6}H_{5}\cdot C_{6}H_{4}\cdot CO\cdot NH\cdot NH_{2}+CI\cdot CPh_{3}\longrightarrow C_{6}H_{5}\cdot C_{6}H_{4}\cdot CO\cdot NH\cdot NH\cdot CPh_{3}\\ I. & II. & \downarrow\\ C_{6}H_{5}\cdot C_{6}H_{4}\cdot CO\cdot CPh_{3} & \longleftarrow\\ IV. & C_{6}H_{5}\cdot C_{6}H_{4}\cdot CO\cdot N=N\cdot CPh_{3}\\ & III. & \downarrow\\ & (warm\ in\ air)\\ & Ph_{3}C\cdot O\cdot O\cdot CPh_{3} \end{array}$$

(e) The acetyl compound, Ph₃C—N=N—CO·CH₃, when heated in presence of air yields traces of diacetyl peroxide, acetic acid and pinacone, which must have been produced by the oxidation and reduction of *free acetyl*, CH₃·CO—, thus:



This work, if not decisive, at least gives a strong indication that a non-ionic mechanism should be assigned to the decomposition of all azo-compounds on heating.

(f) Several of the reactions of the **diazenes** can only be explained by assuming the production of free neutral radicals. Thus diazoacetic ester ¹ and diphenyldiazomethane ² decompose on heating with elimination of nitrogen and union of two divalent residues:

2EtOOC·CH:N₂
$$\rightarrow$$
 2N₃ + EtOOC·CH=CH·COOEt
2Ph₂C:N₂ \rightarrow N₃ + Ph₂C=N-N=CPh₃.

¹ Darapsky, Ber., 1910, 43, 112.

² Staudinger, Ber., 1916, **49**, 1886.

By the action of nitrogen peroxide (which can here be regarded as a free radical) nitrogen is again eliminated from the diazenes, but is replaced by two nitro groups.¹

(i) EtOOC·CH:
$$N_2 + 2NO_3 \rightarrow EtOOC·CH < NO_3 $NO_3 + N_3$$$

(ii) Diazofluorene
$$\begin{array}{c}
C_0H_4 \\
C_0H_4
\end{array}
C:N_2 + 2NO_2 \longrightarrow \begin{array}{c}
C_0H_4 \\
C_0H_4
\end{array}
CNO_2 + NO_2$$

On the other hand, the decomposition of diazoacetic ester by water or alcohol:

EtOOC·CH:
$$N_1 + H$$
·OH \rightarrow EtOOC·CH₁·OH + N_2

and also the Curtius reaction (p. 363) are undoubtedly ionic reactions, since they are catalysed by hydrogen ions.

(g) The final stages of the **diazo-reactions** of Sandmeyer, Pschorr, Bamberger and others may also be of non-ionic type

 $C_6H_5\cdot N_2\cdot OH \longrightarrow C_6H_5\cdot OH + N_2\;;\; C_6H_5\cdot N_2\cdot Cl \xrightarrow{Cu_5Cl_2} C_6H_5\cdot Cl + N_2$ These reactions occur both explosively on heating the dry substances, and, more slowly, in solution. The yield is practically the same whether the reaction is carried out in an acid or in an alkaline solution. The decomposition probably proceeds by the mechanism originally suggested by Kekulé, according to which nitrogen is eliminated from an unstable diazo-compound, containing only trivalent nitrogen, and not from the more stable diazonium salts. Recently Grieve and Hey? have brought forward strong evidence to indicate that the evolution of nitrogen is a fion-ionic change, showing (i) that the rate of the decomposition is practically independent of the solvent employed. and (ii) that aryl radicals are probably set free to react in a similar manner to the decomposition products of aryl azocompounds investigated by Wieland (p. 146).

In the Sandmeyer reaction complex cuprous-diazonium salts are formed, which can often be crystallized from

¹ Wieland, Annalen, 1913, 401, 244.

⁸ J. Chem. Soc., 1934, 1797; cf. J. Chem. Soc., 1934, 1966. Compare Wieland, Annalen, 1934, 514, 145.

reaction-mixtures. These are covalent co-ordination compounds, and may therefore eliminate nitrogen by a nonionic reaction, in such a way that neutral radicals are liberated in close proximity to each other, and are thus able to unite together smoothly. The decompositions of diazo-compounds are, however, always accompanied by production of tar. This complexity of reaction is indeed a frequent characteristic of non-ionic reactions, since the extremely active free radicals are able to decompose in diverse fashions.

Decomposition of Carbonyl Compounds.

Carbon monoxide is isosteric with nitrogen gas, and can also be eliminated from organic compounds in non-ionic manner. Not only will compounds like formaldehyde and formic acid decompose on heating in the absence of catalysts to give carbon monoxide (compare p. 143), but many aldehydes and ketones will decompose photochemically, yielding carbon monoxide and a hydrocarbon 1:

e.g.
$$CH_3$$
— CO — $CH_3 + hv \longrightarrow CH_3$ — $CH_3 + CO$
 C_6H_5 — CO — $H + hv \longrightarrow C_6H_5$ — $H + CO$

Halogen Substitution.

Halogen substitution can either be (i) non-ionic and photochemical or (ii) polar, and dependent on ionic catalysis.

- (i) The photochemical reaction, $H_2 + Cl_2 = 2HCl$, may be regarded as the parent type of many reactions in which the
- ¹ Kirkbride and Norrish, *Trans. Faraday Soc.*, 1931, **27**, 404; Norrish, *ibid.*, 1934, **30**, 107, 119; Frankenburger, *ibid.*, 1931, **27**, 409.

hydrogen of an aliphatic compound is replaced by a halogen under similar conditions, e.g. $CH_4 + Cl_2 = CH_3 - Cl + HCl$. Substitution of one halogen for another under the influence of light is obviously an exactly similar reaction, e.g.

$$CCl_3 \cdot Br + Cl_2 = CCl_4 + HBr$$

This reaction 1 proceeds only slowly in the dark, and at a velocity which is proportional to the concentration of each of the reactant molecules; but it also proceeds quantitatively, in accordance with Einstein's law of photochemical equivalence, under the influence of ultra-violet light. The initial change is probably a photochemical activation of chlorine molecules to atoms, since it proceeds smoothly at a wave-length of 4,100–4,500, whereas light of shorter wave-length, e.g. 3,350 A.U., is required to activate trichlorobromomethane, which then dissociates with separation of bromine,

$$_2CCl_3 \cdot Br = CCl_3 \cdot CCl_3 + Br_2$$

This latter reaction is of course analogous to the non-ionic photochemical decompositions of the halogen acids.

(ii) On the other hand, Lapworth has proved conclusively that halogen-substitution in ketonic compounds proceeds through an intermediate enolic form, and the same mechanism has been extended to explain the α-halogenation of carboxylic acids by the Hell-Volhard reaction.² In general, halogen-substitution in an aromatic nucleus appears to take place under ionic conditions, since these substitutions proceed at low temperatures and in the absence of light, but in the presence of polarizing catalysts, such as iodine chloride, ferric salts or pyridine (see Chapter XVI). A clear picture of the two mechanisms of halogen-substitution is provided by the chlorination of toluene.

(i) Substitution in the side-chain

$$C_6H_5$$
— $CH_8 + Cl_3 = C_6H_5$ — CH_3 — $Cl + HCl$

This can occur under non-ionic conditions, i.e. in the vapour

¹ Noddack, Zeit. Elektrochem., 1921, 27, 359; Rancke, Zeit. Elektrochem., 1921, 27, 365.

⁸ Compare, however, Watson, Chemical Reviews, 1930. 7, 173.

phase, in the presence of light. It also requires as high a temperature as possible.

(ii) Nuclear substitution,

$$C_6H_5-CH_3+Cl_2=Cl\cdot C_6H_4-CH_3+HCl$$

This can proceed in the dark at a low temperature, but must be catalysed by polar substances.

It has already been mentioned that the molecule of

Autoxidation.

oxygen is unique in possessing a magnetic moment corresponding to two Bohr magnetons (p. 121). It can, therefore, be given a formula containing unpaired electrons, :O:O:, and may be classed with free radicals such as nitric oxide. It also possesses the chemical properties of a free radical, since it reacts instantly with nitric oxide, triphenylmethyl and other known free radicals (the halogens being excepted) to form neutral addition compounds.

Molecular oxygen resembles symmetrically-activated ethylene in possessing two adjacent points of free affinity. Peroxide formation is therefore unusually prevalent in oxidations with oxygen gas, and often appears to be a spontaneous reaction, as would be expected from a substance already possessing two free affinities. Spontaneous oxidation, termed autoxidation, has been shown to consist in the addition of a molecule of oxygen 1 to the autoxidized substance, and is an action which can be photochemically activated.2 That autoxidation, even of simple organic substances, is essentially the addition of molecular oxygen to an unsaturated linkage has been demonstrated by investigation of the autoxidation of aldehydes. Thus, when dry oxygen is bubbled at o°, through liquid acetaldehyde which is being illuminated by a mercury lamp, peracetic acid is produced, at a rate depending upon the square root of the intensity of the light.3

¹ Cf. Moureu and Dufraisse, 2nd Solvay Report, 1926, pp. 524-80.

⁸ Cf. Berthoud, 3rd Solvay Report, 1928, pp. 156-65.

⁸ Bowen and Tietz, J. Chem. Soc., 1930, 234.

Benzaldehyde, similarly, is converted by oxygen to benzoic acid, in presence of sunlight, with intermediate formation of perbenzoic acid.¹

Ph·CHO + $O_2 \rightarrow Ph·CO_3H$; Ph·CO₃H + Ph·CHO = 2Ph·CO₂H The first stage of this reaction is very susceptible to both positive and negative catalysis, the positive catalysts being activating surfaces, all of which probably possess a surface film of water, whilst the negative catalysts, or "Antoxygens" as they are often called, are easily oxidizable, and easily absorbed, substances. Suitable antoxygens, such as sulphites, iodine, quinol, or benzyl alcohol, can prevent atmospheric oxidation of an aldehyde when present in a concentration as low as 0.01%.

Reiff ² and Brunner ³ have both shown that at the commencement of direct oxidation of benzaldehyde there is an "induction period" in which oxygen is absorbed but no benzoic acid is formed, followed by a period of very much more rapid reaction, which gradually slackens in velocity, in which benzoic acid is produced. The catalysts markedly alter the duration of the induction period, in which, it has been suggested, there is being formed at an activating surface a sufficient concentration of the molecular oxidation product, Ph CO₂H, to initiate subsequently an oxidation to benzoic acid of the bulk of the benzaldehyde. For the rapid bulk reaction Bäckström has suggested a chain mechanism in which one molecule successively activates another. ⁴

The direct oxidation of olefines certainly, and the direct oxidation of the paraffin hydrocarbons probably, proceed similarly, with the initial formation of a reactive addition-product with a molecule of oxygen. For instance, Brunner and Rideal have shown that the direct oxidation of n-hexane exhibits an induction period analogous to that found in the case of benzaldehyde. The rapid reaction which

¹ Van der Beck, Rec. trav. chim., 1928, 47, 286, 301.

³ J.A.C.S., 1926, 48, 2893. ³ Helv. Chim. Acta, 1927, 10, 767.

British Chemical Abstracts, A, 1927, 1151.

⁵ Stephens, J.A.C.S., 1928, **50**, 568.

⁶ J. Chem. Soc., 1928, 1162, 2824; Helv. Chim. Acta, 1928, 11, 881.

follows the induction period certainly occurs in the bulk phase and has the characteristics of the explosive detonations of petroleum + oxygen mixtures. By analogy with the case of the aldehydes it has been suggested that there is first formed, at isolated centres of activation, an unstable peroxide which is capable of activating the subsequent reaction by a chain process of transmission of energy. The correspondence of the two examples of oxidation is made more complete by the fact that negative catalysts can very markedly prolong the induction period of the oxidation of n-hexane.

The use of lead tetra-ethyl as an "anti-knock" substance in the detonation of petroleum-air mixtures in internal combustion engines is probably to be explained through consideration of the rôle of the intermediate peroxides produced in the combustion process.¹

Moloxides and Molozonides.

The initial addition-product of the reaction between molecular oxygen and an unsaturated compound has been termed a *moloxide*, and will of course be a compound with a sterically unstable four-membered ring. Such a compound may remain unchanged at a low temperature, as, for example, pinene peroxide, which can be kept for years; or it may decompose with explosive violence. The fission of the four-membered ring takes place with formation of a carbonyl compound, and again may be a non-polar dissociation:

e.g.
$$Ph$$

$$C \longrightarrow CH_{2} + \cdot O \longrightarrow O \longrightarrow Ph$$

$$Ph$$

$$C \longrightarrow CH_{2} \longrightarrow Ph$$

$$O \longrightarrow O \longrightarrow CH_{2}$$

Staudinger 2 has studied this type of addition-reaction, mainly with reference to the ketens, which form addition-

¹ Egerton, Nature, 1928, 121, 10.

² "Die Ketene" (1912); Ber., 1925, **58**, 1075; cf. 2nd Solvay Report, 1926, pp. 593-9.

compounds with oxygen at low temperatures. These are white powders, which are liable to explode violently and spontaneously, yielding ketones and carbon dioxide.

Their great reactivity, and ready polymerization to four-membered ring compounds (which latter can dissociate again), have led Wieland 1 and others to classify the ketens themselves as *diradicals* to which formulæ such as (CH₃)₂C——C——O may be assigned; but such a con-

clusion is not justified unless supported, for instance, by measurements of the parachor, since the exceptional reactivity of these compounds could be attributed equally well to an exceptional strain in the system of contiguous double bonds. The ready addition of ketens to such unsaturated compounds as benzoquinone, azobenzene, pyridine and nitrosobenzene are also quite exceptional reactions, which recall those of the aliphatic diazo-compounds.

The moloxides have been isolated only as complex polymers of the initial labile four-membered ring-compounds. Ozone is even more ready than oxygen to form addition-compounds which are described as *molozonides*. Thus, unlike oxygen, it unites directly with olefines; but the products, when stable enough to be isolated, are again found to be polymers of high molecular weight. Staudinger assigns formulæ of the following types to these polymers:

¹ Annalen, 1913, 401, 243.

These formulæ serve to explain why aldehydes or ketones are produced, as well as carboxylic acids and hydrogen peroxide, when the ozonides are decomposed with water; but, since the fourth valency in quadrivalent oxygen is necessarily an electrovalence, the ozonides would now be formulated with a semipolar bond, as follows:

Influence of Solvent on Non-ionic Reactions.

The influence of solvents on the velocity of two typical non-polar photochemical reactions is set out in the following table.

TABLE III.

Relative Reaction Velocities of Photochemical Changes

Solvent.			xidation of odoform.1	Polymerization of Anthracene.
Carbon tetrachlorie	de		39.4	
Benzene			34.6	355
Toluene			-	163
Xylene				52
Ether			17.0	
Carbon disulphide			24·I	
Anisole				18.6
Phenetole .			*****	18.0
Ethyl acetate.			8.1	
Methyl alcohol			3.2	
Ethyl alcohol.			2.3	andahiya
Acetone			0.7	

These numbers are in harmony with the anticipation that polar solvents, such as water or acetone, would induce electrical polarizations in reactant molecules, rendering them less symmetrical and thus diminishing the chances of symmetrical non-polar activation.

In contrast, a typical polar action, like the isomerization of phenyl allyl acetate (page 319), occurs more rapidly in the more polar solvents. This is also true of the ionic reactions discussed in the following chapter.

¹ Plotnikow, Zeit. phys. Chem., 1911, 75, 396.

Luther and Weigert, Zeit. phys. Chem., 1905, 53, 393.

Inhibition of Photochemical Reactions.

A number of substances are known, of which oxygen and oxides of nitrogen are typical, which behave as powerful inhibitors to photochemical reactions. They are believed to act by combination with the free active atoms or free radicals, in such a way as to terminate a chain-reaction. F. Perrin has noted that the substances which inhibit photochemical reactions also inhibit fluorescence, and has pointed out that the inhibitors in both liquid and gaseous systems are all substances which can easily lose an electron. Thus, the inhibitors may be either (i) free radicals, or similar substances containing an odd electron, e.g. nitric oxide, which can combine with free radicals to form photochemically inactive products, such as NOCl, or (ii) substances, such as potassium iodide, which by interchange of an electron can give free radicals of a more stable type than those originally produced by the absorbed light, e.g.

(i)
$$Cl_2 + h\nu \longrightarrow 2Cl$$
·
(ii) $Cl_1 + I \longrightarrow Cl_1 + I$

Thus the free chlorine atom has an electron-affinity of + 116,000 cal., which is available for propagating a chain-reaction; but the iodine atom, which is formed by the transfer of an electron from an iodide ion to a chlorine atom, has an electron-affinity of only + 81,000 cal. and may not be able to activate the same chain process.

Inhibition in liquid media becomes less efficient in the more viscous solutions, in accordance with Perrin's conclusion that the inhibiting substance must react on collision with the photochemically activated radical.

Conditions Favouring the Existence of Free Radicals.

The breakdown of the dogma of the constant quadrivalency of carbon by the discovery of triphenylmethyl was originally explained by stereochemical considerations. Thus, it was supposed that triphenylmethyl existed as a

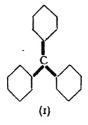
¹ Ann. de physique, 1929, 12, 169-275.

free radical because the spatial requirements of the large phenyl groups were such as to render impossible the packing of more than three of these groups around a single carbon atom. This conclusion, however, was disproved by Gomberg's preparation of the stable hydrocarbon tetraphenyl methane by the following reaction:

$$Ph_{\bullet}C \cdot Br + Ph \cdot Mg \cdot Br = Ph_{\bullet}C + MgBr_{\bullet}.$$

The yield was only of the order of 5%, but, since triphenylmethyl chloride will act *quantitatively* with either phenol or aniline to give products such as Ph₃C·C₆H₄·OH or Ph₃C·C₆H₄·NH₂, it is obvious that steric considerations alone are of little value.

Flürscheim, in 1905, applied to the problem of tervalent carbon the theory of variable chemical affinity-demand (p. 429), which he had developed from Werner's conception of valency. His formula (1) for triphenylmethyl was intended to indicate that the large "affinity requirements" of the three phenyl groups left the tervalent carbon atom with



an insufficient residuum of affinity to form a stable bond with a fourth group. This theory explained successfully the dissociation of the "carbonium salts"; but its general application would indicate the dissociation of many carbon compounds which are particularly inert, e.g. carbon tetrachloride should tend to dissociate into CCl₂ and Cl on account of the large affinity-requirements of carbon-chlorine links. Conversely, tetraphenylhydrazine should not tend to dissociate, since tervalent nitrogen possesses a large reserve of residual affinity. The behaviour of derivatives of triphenylmethyl is also not always in accordance with

the anticipations of Flürscheim's theory. Thus Schlenk and Mark 1 have shown that decaphenylbutane (II) is incapable of existence, since it dissociates immediately and completely into pentaphenylethyl, Ph₂C—CPh₂, by the rupture of a link which should be particularly strong.

On the other hand, Burton and Ingold ² have pointed out that the dissociation of organic substances into stable free radicals occurs only in compounds containing aromatic rings, or similar tautomerically-active unsaturated groups, attached directly to the atom at which the dissociation occurs. They therefore suggest that "The power of the phenyl group to compensate electrostatic disturbances of either sign in an attached atom is doubtless the cause of the dissociability of the hexa-arylmethyls and allied compounds." Polynuclear aromatic groups act more powerfully than the simple phenyl group in stabilizing free radicals of the triarylmethyl type; similar considerations are found to hold for the stability of the metallic ketyls and their analogues.³

E. Hückel 4 has investigated the problem from the theoretical standpoint, and has shown that interactions resulting from the mutual perturbations of electronic orbits will tend to stabilize a free radical such as Ph₂C, to which a symmetrical "resonance structure" (p. 396) should be assigned. Further, the covalent bonds of a compound like hexaphenylethane are subject to a greater internal "elec-

¹ Ber., 1922, 55, 2285.

² J. Chem. Soc., 1928, 904; Proc. Leeds Phil. Soc., 1929, 1 (iv), 421. ³ Schlenk and Bergmann, Annalen, 1928, 463, 1; Ziegler and Boye, Ann., 1927, 458, 248; Ingold, Annual Reports, 1928, pp. 150-56; Bennett and Chapman, Annual Reports, 1930, pp. 125-9.

⁴ Zeit. für Physik., 1933, 83, 632; Trans. Faraday Soc., 1934, 30, 40, 59.

158 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

trostatic bond strain" than are the bonds in ethane, and hence less energy will be required to bring about dissociation into radicals. Hückel's theory is the mathematical equivalent of Ingold's conception.

¹ Cf. Trans. Faraday Soc., 1934, **30**, 52; J. Chem. Soc., 1933, 1120; Chemical Reviews, 1934, **15**, 225.

CHAPTER VIII

IONIZATION AND IONIC REACTIONS

Introduction.

Berzelius's dualistic theory postulated that molecules were formed by the union of oppositely-charged radicals into which they could often be resolved by electrolysis. His extension of this theory to the field of organic chemistry was based on the assumption that carbon-compounds, although they were not electrolytes, were nevertheless also built up from oppositely-charged radicals. This theory implies that organic reactions depend on an interchange of ions, just like the double-decomposition of a pair of salts. On the other hand, Dumas's discovery of substitution and the subsequent development of structural chemistry on a strictly non-polar basis, led inevitably to the adoption of schemes in which all organic reactions, except those of organic acids, bases and salts, were formulated as depending on an interchange of neutral radicals.

The latter view remained unchallenged until the end of the nineteenth century, when Euler in 1899 1 put forward a theory of catalytic reactions, according to which the hydrolysis of an ester was attributed to its dissociation into ions. He attributed the accelerating action of catalysts to an increase in the velocity of ionization of the ester, although the equilibrium-constant of this reversible process was not supposed to undergo any important change by reason of the presence of the catalyst. In general, he considered that there were no grounds for making a distinction between electrolytes and non-electrolytes, and that all

¹ Öfvers. K. Vet. Akad. Stockholm, 1899, 56, 369 (Abstr., 1900, ii, 532); Ber., 1900, 33, 3202; Zeit. phys. Chem., 1901, 36, 641.

chemical substances might therefore be regarded as electrolytes.

A similar suggestion was made by Lapworth, who expressed the view ¹ that it is to electrolytic dissociation, often doubtless in extremely minute amount, that the majority of changes in organic compounds may be most probably assigned. This mechanism was used to provide an explanation of the "Intramolecular Changes in Organic Compounds", which he had classified with marked success in 1898 ² into the $\alpha\beta$ - and $\alpha\gamma$ - types. In particular, it was suggested that "The ions of organic compounds usually possess only one point of free valency", which is merely another way of saying that they are usually weak electrolytes, and therefore form only univalent ions.

During the first decade of the present century, electrical theories of organic reactions were put forward in America by Michael, Stieglitz, Fry and others 3; but no really satisfactory theory was possible until G. N. Lewis in 1916 demonstrated the existence of two types of valency, which Langmuir in 1919 distinguished as covalence and electrovalence. It then became clear that a bond could be disrupted in two different ways, giving rise either to two neutral radicals or to two oppositely-charged ions. The plus and minus signs, which the American workers had superposed on the bonds of the ordinary structural formulæ, could then be interpreted as indicating the sign of the electrical charge which a given atom might be expected to carry when ionized. Since, however, these signs have been used constantly to represent the actual charges of the ions, e.g. in ionic aggregates such as NaCl, it was obviously undesirable to use the same symbol to indicate a mere tendency to ionize in a particular way. Thus the addition of positive and negative signs to the formulæ of compounds such as methyl iodide is only justified when

¹ J. Chem. Soc., 1901, 79, 1266.

² J. Chem. Soc., 1898, 73, 445.

⁸ Cf. Henrich, "Theories of Organic Chemistry", 1924 translation.

ionization is postulated as having actually taken place, as in the scheme

$$CH_{2}I \rightleftharpoons \overset{+}{C}H_{2} + \overset{-}{I}.$$

For this reason, in a discussion of "The Polarity of Double Bonds", pairs of plus and minus signs were used by Lowry to replace bonds which had been ionized, instead of being superposed on the unbroken bonds, as in the earlier formulæ of Fry and others. Thus the activation of an aldehyde or ketone under the influence of a polar environment was represented by the scheme $>C \longrightarrow >C \longrightarrow >C \longrightarrow >C \longrightarrow >C$, where the symbol + or - represents a deficit or excess of one planetary electron in the shell of an atom as compared with the positive charge on the nucleus. Since, however, ionization is necessarily preceded by the development of a local dipole moment, all risk of confusion can be avoided by using the differential symbols $\delta +$ and $\delta - 2$ to represent a tendency to ionize in a particular manner, as indicated by an electrical displacement in the same direction, e.g.

$${}^{\delta +}_{CH_3} \stackrel{\delta -}{I} \rightleftharpoons {}^{+}_{CH_3} + \stackrel{-}{I}_{Or} > {}^{C=O}_{C} \rightleftharpoons > {}^{+}_{C} - \stackrel{-}{O}_{C}$$

The introduction of + and — signs as a means of expressing Lapworth's ionic theory of organic reactions was followed by ten years of intense controversy; but a consistent electrochemical theory has at last emerged, which takes into account the fact that organic compounds (as well as mineral salts) are composed of nuclei and electrons. Compounds built up in this way may therefore possess dipole moments and under appropriate conditions may even undergo an ionic type of disruption. This ionic type of disruption, however, is now seen to be characteristic mainly of reactions which take place in solution and under the influence of ionized reagents or catalysts, since photochemical reactions, and thermal reactions in gases, have already been

¹ J. Chem. Soc., 1923, 123, 822.

² C. K. and E. H. Ingold, J. Chem. Soc., 1926, 1312 (footnote); Chemical Society Annual Reports, 1926, p. 135.

shown in the preceding chapter to be predominantly nonionic in character. The present chapter, therefore, contains a summary of the evidence from which the ionic character of certain specific types of organic reactions may be inferred.

Ionization of the Reagents of Organic Chemistry.

Many of the reagents used in organic chemistry, such as the mineral acids and the caustic alkalis, are inorganic compounds, which are already ionized, either completely or reversibly, and owe their activity to the ions which they can provide. Other inorganic reagents, such as the metals and the halogens, are converted into ions during the course of the reactions in which they are used. A large number of highly reactive organic compounds, which are not electrolytes, are, however, also used as reagents. Some of these, such as the acyl and alkyl halides and dimethyl sulphate, on the one hand, and the organo-metallic compounds on the other, yield an inorganic anion or kation (together with a covalent organic compound) during the course of the reaction; but others, such as the alcohols, aldehydes and ketones, aniline and phenylhydrazine, yield only water as an inorganic product of the interaction of two organic compounds. In some of these cases it is quite clear that the reaction is ionic in character, since the activity of the reagent is exerted through one or both of its constituent ions; but in others the validity of this conception is less obvious, and must form the subject of inquiry before it is accepted in any particular case. In the following paragraphs, some of the principal reagents used in organic chemistry are considered from this point of view.

(a) Alkalis and Metallic Salts. Metallic salts as a class may be regarded as permanently ionized, since there is no shortage of electrons to justify a demand for electron sharing, and any attempt to share electrons would therefore violate the octet rule. The hypothesis of complete ionization applies not only to metallic salts like Na Cl, but

also to the alkalis, and to organic derivatives such as sodium ethoxide, Na OEt, sodium acetate, NaO·CO·CH₃, sodio-malonic ester, Na CH(COOEt)₂, etc., in all of which a permanent ionization of the metallic radical is generally admitted. Since one of the ions persists in many of the purely organic reactions of these substances, the view that the mechanism of the reaction is ionic can scarcely be disputed.

(b) Metallic Alkyls. Although metallic salts are generally formulated as mere aggregates of ions, this formulation is no longer valid in the case of co-ordination-compounds, where the metal may be linked to the co-ordinated radicals by dative bonds. It is, therefore, not safe to assume that a metallic derivative of an organic compound is necessarily ionized, unless it has been shown experimentally to behave as a strong electrolyte in an ionizing solvent which does not react chemically with it. This consideration is particularly important when the physical properties of the compound in question (e.g. volatility and solubility in organic solvents) resemble those of a typical organic compound rather than those of a metallic salt.

On the other hand, even if the metal of an organo-metallic compound is initially attached by covalent bonds to the associated radicals, the rupture of these bonds to form a metallic kation is obviously of the unsymmetrical type, which leads to the formation of ions and not of neutral radicals. This rupture can indeed be regarded as a reversal of the process of co-ordination, in which the covalences of the dative bonds are reconverted into electrovalences. Thus the intense reactivity of the metallic alkyls may be attributed to a permanent ionization of the reagent as shown in the following formulæ:

(1)
$$\stackrel{+}{\text{Na}} \stackrel{-}{\text{CH}}_{1} \cdot \text{C}_{6}\text{H}_{5}$$
 $\stackrel{-}{\text{CH}}_{3}\stackrel{+}{\text{ZnCH}}_{3}$;

To the free organic anion, too, must be ascribed a great readiness to undergo chemical change 1 as a complex mix-

¹ Carothers and Coffmann, J.A.C.S., 1930, 52, 1254.

ture of decomposition products results when the sodium alkyls are heated.

Since zinc methyl is a volatile liquid, with none of the physical properties of a salt, it can also be formulated as a covalent compound, like mercuric chloride or the mercury alkyls. Its exceptional reactivity may then be attributed to the fact that the outer shell of the atom of zinc carries only four electrons:

Similarly, the alkyl-derivatives of quadrivalent lead, tin, and germanium, might be formulated as salts; but, since they are volatile liquids, like the hydrocarbons, it is more probable that they are covalent compounds like methane and silicon hydride. Their greater stability is in harmony with the fact that the central atom in these compounds would then carry a complete octet of electrons. The thermal decomposition of lead methyl and lead ethyl, into atomic lead and the free neutral alkyl radicals, has already been cited as an example of a non-ionic reaction. Mercury diethyl appears to act similarly, but this does not exclude the possibility of an interchange of ions if these compounds were used under polar conditions.

(c) The Grignard Reagent. The Grignard reagents require separate consideration, since, in the normal laboratory procedure, the magnesium alkyl halide separates with two molecules of ether of crystallization. Mills has therefore suggested that the complex should be regarded as a co-ordination-compound (IV) or (VI). The formation of this co-ordination-compound from the ionic aggregate (V) implies that the bivalent magnesium atom, which had already lost its two outer electrons in forming an ion, acquires a complete octet of electrons by sharing four pairs with other atoms. In this process it also acquires a double

¹ Taylor and Jones, J.A.C.S., 1930, **52**, 1111.

negative charge, whilst the tervalent oxygens are formulated with a positive charge as in OH_a .

Whilst, however, ether may play an important part in the preparation of the Grignard reagents, it does not appear to take any direct part in any of their reactions, since the simple ionic aggregates (V), although much more difficult to prepare, have exactly similar chemical properties. We may therefore attribute the activity of the Grignard reagents, if not to the presence of free ions, at least to the ease with which they can be converted into ionic products. This is substantiated by the fact that in ordinary ethereal solutions equilibria always occur in accordance with the scheme

$$2MgR hal \rightleftharpoons MgR + Mg hal$$

(d) The Acids. The most obvious property of the acidic hydrides is the readiness with which they can be ionized, either by interaction with water, which converts them into "oxonium" salts, or by neutralization with nitrogenous bases, which converts them into "ammonium salts":

$$HCl + OH2 = OH3 + Cl$$

$$HCl + NH3 = NH4 + Cl.$$

It may therefore be inferred that nearly every action of any acid, occurring at moderate temperatures and in the absence of photochemical activation, depends upon a process of ionization. It is, however, necessary to point out that organic chemists have been accustomed to treat the strongest mineral acids as exerting the characteristic reaction of an alkali in equations such as

$$\begin{array}{ll} C_{6}H_{6} + HO \cdot NO_{8} &= C_{6}H_{5} \cdot NO_{9} &+ H_{9}O \\ C_{6}H_{6} + HO \cdot SO_{9} \cdot OH &= C_{6}H_{5} \cdot SO_{8} \cdot OH &+ H_{9}O. \end{array}$$

¹ W. Schlenk and W. Schlenk, jun., Ber., 1929, **62**, 920; Gilman and Fothergill, J.A.C.S., 1929, **51**, 3149.

These equations represent the raw materials and the finished products of chemical actions, but they do not harmonize with the fact that the reagents in question are strong mineral acids, unless we are prepared to admit that all oxyacids may, like water, be regarded as potentially amphoteric substances. This possibility has, however, been formulated in a very plausible form by E. E. Turner, who has suggested that "orthonitric acid" and its sulphate may actually yield hydroxyl ions as in

$$^{+}$$
OH $^{-}$ OH and $[N(SO_4H)_4]^+$ OH-

Moreover, investigation of the absorption spectrum and of the electrical conductivity of concentrated nitric acid indicates that it possesses the covalent structure HO—NO₃, containing the same nitro-group as non-ionic compounds such as nitrobenzene, C₆H₈—NO₃. Dilute nitric acid, on the other hand, gives a spectrum resembling that of the nitrates, such as K NO₃, which contain the nitrate anion, NO₃.²

(e) Alkyl Halides and Sulphates. It is a noteworthy fact that the alkyl sulphates and iodides are active alkylating agents, whereas the alkyl esters of carboxylic acids are not. This contrast is probably due to the greater strength of the mineral acid, i.e. to its greater tendency to ionization. The ionization of an acid enables it to get rid of its hydrogen atoms in the form of positively-charged protons, and thus in some measure to satisfy the electron-hunger of the "negative" radicals, which seek to pass into anions with one or more surplus negative charges. Even this process, however, can only occur in presence of water or of some other substance which is sufficiently basic to accept the protons from the acid.

The ionization of an alkyl sulphate does not take place with the same ease, or under the same conditions, as that

¹ J. Chem. Soc., 1928, 337. ² Hantzsch, Ber., 1925, 58, 941.

of an acid, since it is much more difficult to convert the alkyl groups into ions; but in presence of a base, or of an organic compound which can accept an alkyl group, the transfer can occur. Since, however, this transfer is accompanied by the formation of a negatively-charged inorganic ion, it is reasonable to suppose that at the moment of transfer the alkyl group carried a positive charge, and reacted in the form of a kation. The slowness of the interaction of a base with an alkyl halide, as contrasted with the velocity of neutralization of a base with hydriodic acid, may be regarded as a measure of the greater difficulty of ionizing the alkyl iodide; and in the same way the even smaller velocity of interaction with alkyl bromides or chlorides (cf. Chapter X) may be regarded as a measure of the still greater difficulty of ionizing these compounds.

(f) Halogens. The bond between the atoms in the molecule of any halogen must be, at least statistically, symmetrical. This symmetry is maintained when the molecule is resolved into neutral atoms by photochemical activation, as described in the previous chapter; but it has been proved experimentally that in the absence of light the reaction of chlorine or bromine with ethylene proceeds preferentially under unsymmetrical polar conditions (p. 172).

Unsymmetrical fission of a halogen molecule results in the breaking down of the covalent bond into a pair of oppositely-charged ions:

$$CI$$
— $CI \rightleftharpoons \overset{+}{C}I + \overset{-}{C}I$; or $\overset{-}{C}I : \overset{-}{C}I : \rightleftharpoons : \overset{-}{C}I + : \overset{-}{C}I$:

This theory of the amphoteric character of the halogens has been advocated for many years by W. A. Noyes, who has done much work in order to establish the existence of the positively- (as well as of the negatively-) charged ions of these elements. Thus the reversible action of the halogens with water

$$Cl_2 + H_2O \rightleftharpoons HCl + HClO$$

¹ See, for instance, J.A.C.S., 1901, 23, 460.

can be formulated most readily as depending on an interchange of ions:

$$Cl_1 + H_2O \rightleftharpoons \stackrel{+}{C}l\stackrel{-}{C}l + \stackrel{+}{H}\stackrel{-}{O}H \rightleftharpoons H--Cl + Cl--OH.$$

Conversely, the decomposition of hypochlorous acid into Cl₂O and H₂O can be represented most easily as depending upon a dissociation of this very weak acid according to the two methods set out below, one of which again involves the formation of positively-charged chlorine ions.

$$HO-Cl \rightleftharpoons H \stackrel{-}{O} + \stackrel{+}{Cl}$$
 $H-OCl \rightleftharpoons H \stackrel{-}{O} + \stackrel{-}{Cl}$
 $H-OH + Cl_{3O}$

Lewis ¹ finds physical support for the same view in the fact that liquid iodine possesses an appreciable electrical conductivity, K = 2.16 to 6.8×10^{-5} at 140°, a result which indicates that the halogen must be ionized to some extent

according to the equation $I_1\rightleftharpoons I+I$. Confirmatory evidence is afforded by the fact that the halogens also behave as feeble electrolytic conductors when dissolved in liquid sulphur dioxide.² In the case of iodine, the liberation of the element at the kathode has been recorded as occurring in the electrolysis of fused IBr and ICl, and of ICl₂ in CHCl₂·COOH. Moreover, although the iodide ion is always negatively charged, positively-charged ions of iodine are found in kathode rays.

Further evidence of the existence of "positive" halogens can be obtained from the study of organic halogen compounds (Chapter X) and most especially of those which act as oxidizing agents and are prone to "inverse substitution"

The Mechanism of Ionic Reactions in Organic Chemistry.

Ionic reactions in organic chemistry generally include three distinct operations, (i) ionization of the reagent, if

¹ J.A.C.S., 1916, **38**, 762.

² Walden, Zeit. phys. Chem., 1903, 43, 407; Henderson, Hirst and Macbeth, J. Chem. Soc., 1923, 123, 1134.

this is not already ionized, (ii) ionization of the reactant, or alternatively the ionization of one link of a double bond, (iii) neutralization of these two pairs of ions with formation of one or two covalent molecules. These operations need not, however, proceed independently and consecutively, but may be coupled together (like a battery and an electrolytic cell), so that the energy liberated by one of the operations is available to meet the requirements of another, and the rôle of the solvent in completing the electrical circuit is undoubtedly very important (see pp. 95–98).

(a) An ionic reaction in which both molecules are ionized and both pairs of ions are then neutralized is provided by the combination of acetaldehyde with ammonia to form aldehyde-ammonia,

$$CH_3 \cdot CH = O + NH_3 \rightleftharpoons [CH_3 \cdot \dot{C}H - \ddot{O} + \ddot{H}NH_3] \rightleftharpoons CH_3 \cdot CH$$

$$NH_3$$

When, however, tetramethylammonium iodide is formed from trimethylamine and methyl iodide, only one molecule of the interacting pair is ionized, and the opposite charges resulting from the ionization persist in the product:

$$NMe_3 + CH_3I \rightleftharpoons [NMe_3 + CH_3\overline{I}] \rightleftharpoons NMe_4\overline{I}.$$

(b) The formation of the tetramethylammonium ion by the addition of CH_3 to trimethylamine is characteristic of a large number of organic reactions, in which an ion attaches itself to an organic molecule. In this process, a covalent bond is formed by the transfer of an electron, either from the organic molecule if the ion is positively charged, or to the molecule if the ion is negatively charged. These actions are therefore classified primarily according to the sign of the charge on the ion (see below, p. 180); but the atom to which the radical of the ion is attached does not necessarily also retain the ionic charge, since, in unsaturated and conjugated systems, this may be handed on to a more

¹ Cf. Robinson, 4th Solvay Report, 1931; Institute of Chemistry Lectures, 1932, p. 18.

distant atom. Thus, in the methylation of aminocrotonic ester, the methyl group attaches itself to the β -carbon atom; but its positive charge is transferred to a distant amino-group

$$NH_3$$
: CMe=CH·COOEt + $\stackrel{+}{C}H_3I \rightarrow I + \stackrel{+}{N}H_3$ =CMe·CH COOEt.

This transfer depends on a migration of electrons, which has been described as an electrotropic change (p. 388).

(c) Lapworth's proof (p. 104) that the cyanhydrin condensation of aldehydes and ketones depends on the presence of cyanide ions has already been cited as evidence of the ionic character of some addition-reactions of unsaturated substances. This reaction is expressed by the following scheme:

$$>$$
C \longrightarrow O + $\overline{C}N \longrightarrow >$ C \longrightarrow O

It provides an interesting contrast to the methiodide condensation described in the preceding paragraph, since, (i) the reactive ion carries a negative instead of a positive charge, (ii) this charge is transferred to an adjacent atom of oxygen instead of to the atom of carbon to which the cyanide ion becomes attached, i.e. from an α - to a β -position in the inolecule. We have, however, already seen that in aminocrotonic ester the charge of the ion can be transferred to a γ -position relatively to the α -position which accepts the ion; similarly in conjugated systems (Chapter XV) the negative charge of a cyanide ion may be transferred to a still more distant atom.

Ionic Additions to Olefines.

(a) Addition of Acids. We have seen that the cyanhydrin ion, formed by condensing the cyanide anion with an aldehyde or a ketone, can be converted into a neutral molecule by the addition of a proton derived from a second-

¹ Hamilton and Robinson, J. Chem. Soc., 1916, 109, 1029.

ary reagent such as potassium bicarbonate. Very many other neutral addition-compounds are formed in this same general way. Thus, at low temperatures liquid ethylene forms an unstable addition compound with hydrogen bromide, which is not identical with ethyl bromide, since it dissociates on melting.¹ The structure of this unstable intermediate-compound is not known, but two alternatives may be suggested. According to one of these the bromide ion, with its lone pair of electrons, might attach itself first to the olefine, in which case the hydrogen of the reagent would retain a measure of freedom in the initial stage:

$$C_2H_4 + HBr \rightleftharpoons (C_2H_4 \cdot Br)^{-1}H;$$

but this type of action is unlikely to occur in view of the great stability and the very small reactivity of the bromide ion. Alternatively, the ethylene may perhaps act as a weak base, in which case the hydrogen of the acid would unite first with the olefine, whilst the anion would remain free, thus:

$$C_2H_4 + HBr \rightleftharpoons (C_2H_4 \cdot H)^+ Br^-$$

The latter alternative, which finds expression in Robinson's classification of ethylene as a "kationoid reagent" (p. 181), has the advantage of recalling in almost every detail the conditions under which a dissociable hydrochloride is formed from methyl ether,

$$(CH_3)_3O + HCl \rightarrow (CH_3)_3OH + Cl.$$

It also finds support in observations on the addition of hydrochloric acid to olefinic terpenes such as pinene and camphene. Both hydrocarbons yield addition-products which are prone to undergo isomeric changes (p. 350); but Meerwein ² has shown that the substances which undergo isomeric change are not the neutral molecules represented by the formula $C_{10}H_{17}Cl$, but positive ions, $(C_{10}H_{10}\cdot H)^+$, derived from intermediate salt-like addition-compounds of the type $(C_{10}H_{16}\cdot H)^+$ Cl^- .

¹ Maass and Wright, J.A.C.S., 1924, 46, 2664.

² Annalen, 1924, 435, 214; 1927, 453, 16.

Kharasch and Mayo ¹ have demonstrated the important part that can be played by catalysts in the reaction between an olefine and hydrogen bromide. In the absence of peroxides, which readily form when an olefine is exposed to the air, an entirely different addition-reaction occurs to that which preponderates when the addition is carried out in an open vessel, or in a solution not containing any anti-oxidants. In view of the characteristic non-ionic reactions of oxygen and the peroxides (pp. 150–153) this discovery is of great interest, for it possibly indicates that ionic and non-ionic addition to an olefine are alternative reactions, each having well-defined characteristics.

- (b) Addition of Halogens. The ionic mechanism suggested above for the addition of hydrogen bromide to an olefine can also be applied to the addition of bromine. Thus the bromination of ethylene can be pictured as a two-stage ionic process, in which the intermediate compound is formulated as $(C_2H_4\cdot Br)^+Br^-$. This view is supported by strong experimental evidence drawn from three distinct sources.
- (i) Stewart and Edlund,² in 1923, showed that the interaction of ethylene and bromine takes place mainly, if not entirely, at the surface of the glass vessel in which the two components are brought together. Since no evidence of a gaseous reaction could be obtained, it is clear that the union of ethylene with a halogen is a catalytic process, which is not spontaneous in the absence of a solid surface. If, however, it were only necessary that the two components should be brought together at an absorbing surface, it would be expected that paraffin wax would be an efficient catalyst. Actually, Norrish ³ found that the action was stopped almost completely by coating the surface of the glass with paraffin wax, and halved by coating it with cetyl alcohol; on the other hand, the velocity was nearly

¹ J.A.C.S., 1933, **55**, 2468, 2521, 2531; 1934, **56**, 712; 1934, **56**, 1212.

² J.A.C.S., 1923, **45**, 1014.

³ J. Chem. Soc., 1923, 123, 3006; cf. J. Chem. Soc., 1926, 55.

doubled by coating the glass with stearic acid. In four series of experiments the average velocity constants for the different surfaces were as follows:

Stearic acid				o∙o86
Glass .			•	0.051
Cetyl alcohol				0.027
Paraffin wax				0.003

Moreover the slow reaction recorded above for a vessel lined with paraffin wax was actually preceded by a "period of induction" during which no action whatever appeared to take place. These experiments, therefore, show clearly that a *polar* catalyst is needed to promote the interaction of ethylene and bromine, and lend strong support to the view that the whole process is ionic in character.

(ii) Gwyn Williams 1 confirmed the conclusions of Stewart and Edlund and of Norrish, that the interaction of ethylene and bromine is a surface reaction, but found that the "order" of the reaction (0, 1 or 2) depended on the condition of the glass surface. Indications were, however, obtained that a reaction which had started at the surface might be propagated as a chain-reaction in the bulk phase; e.g. the actual velocity of reaction was reduced by one-third by packing the glass vessel with pieces of "poisoned" glass-tubing. For a given surface and order of reaction, the velocity coefficient was independent of the initial pressure of ethylene, but rose with initial pressure of bromine.

The rate of reaction decreased with rising temperature, thus suggesting that the action was not a direct combination of the two components, but depended on the formation of an intermediate product, which was less stable at high temperatures.

The surface-action was accelerated by the presence of water, and this effect persisted in a second run, but was destroyed when the adsorbed water was removed by "baking out" the glass vessel at 350-400° and under a pressure

¹ J. Chem. Soc., 1932, 1747 and 1758.

174 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

less than 10⁻⁵ mm. The presence of water also tended to convert a second-order reaction

$$d[C_2H_4Br_2]/dt = k_2[C_2H_4][Br_2]$$

into a first-order reaction

$$d[C_3H_4Br_2]dt = k_1[C_3H_4]$$

which was independent of the concentration of the bromine. It therefore appeared that the intermediate product, which had already been postulated to account for the negative temperature-coefficient, was probably a hydrated bromine molecule, which was decomposed by raising the temperature, but was formed so readily on the glass surface that its concentration was almost independent of the bromine-concentration in the gas-phase.

The precise structure of the hydrated bromine could not be determined, and is not of great importance; but the compound is clearly of a polar character, whether it be formulated as an oxonium salt I or as a mere adsorptioncomplex II.

(I)
$$\begin{array}{c} H \\ + \\ Br \\ - \\ Br \end{array}$$
 (II) $\begin{array}{c} \delta^+ \\ H \\ \delta^- \\ \delta^+ \\ Br - Br \end{array}$

In either case, it is probably the positive ion or the positive pole that is reactive, although the negative ion also unites finally with the olefine just as in the case of a halogen hydride.

(iii) Ethylene and chlorine react in presence of steam to form ethylene chlorohydrin; large percentages of chloro- or bromo-hydrins ¹ are also formed when olefinic substances interact with chlorine or bromine in aqueous solution. On this basis A. W. Francis ² has shown that, under conditions favouring ionization, the initial stage of halogen-addition must be the condensation with the organic molecule of the positive ion of the halogen, which is of high energy content

¹ J. Read and others, J. Chem. Soc., 1917, 111, 240; 1920, 117, 1214; 1921, 119, 1774; 1922, 121, 989; 1928, 1487.

² J.A.C.S., 1925, 47, 2340.

(p. 222). The formation of a bromohydrin, for example, is *not* to be formulated as:

Stage 1 . . .
$$Br_1 + H_2O = HBr + HOBr$$

Stage 2 . . $C_2H_4 + HOBr = Br \cdot CH_2 \cdot CH_2 \cdot OH$

because changes of experimental conditions, which would alter the equilibrium attained in stage I (e.g. working in acid solution and so altering the concentration of HOBr available for addition), do not affect the rate of formation of bromohydrin. Furthermore, hypobromous acid itself is less reactive than bromine water for the production of the bromohydrins.

Since a mixture of dibromide and bromohydrin is invariably produced by the action of bromine water, the active agent must be one constant constituent of the reaction mixture, and this can scarcely be anything other than positively-ionized bromine, Br^+ :

The production of ethylene dibromide may consequently be formulated as follows:

Stage r = Activation .
$$C_2H_4 \rightarrow CH_2 - CH_2$$
 : $Br_2 \rightarrow Br + Br$
Stage 2 = Reaction of one ion . $Br + CH_2 - CH_2$ $Br \cdot CH_2 - CH_2$
Stage 3 = Reaction of second ion $Br \cdot CH_2 - CH_2 + Br \rightarrow Br \cdot CH_2 - CH_2 \cdot Br$

The negative bromide ion is perfectly stable in aqueous solution, and is indeed less liable to lose its ionic charge than the hydroxyl ion, which is necessarily present in small concentration in the water; this hydroxyl ion may therefore take the place of the bromide ion in the final stage of the addition process, and thus give rise to ethylene bromohydrin instead of ethylene dibromide

Stage 3'. . . Br·CH₂—CH₂ + OH → Br·CH₂—CH₃·OH
This view was supported by the observation of Terry and
Eichelberger 1 that, when the concentration of bromide
ions was increased by adding sodium bromide to the
bromine water, the percentage yield of bromohydrin was
decreased relatively to that of the dibromide; moreover,

by adding chloride ions, in the form of sodium chloride, some chloro-bromide could be obtained. Francis confirmed these results, and also showed that no ethylene dichloride is formed by the interaction of ethylene and bromine in presence of sodium chloride, thus proving that chlorine is not a constituent of the reaction mixture; but, in view of the possibility that bromine might have reacted with the sodium chloride to produce some Br—Cl as the additive reagent, he showed that even iodine (which certainly cannot displace chlorine from a chloride) would react with ethylene and sodium chloride to give I·CH₂—CH₂·Cl. Furthermore, in presence of sodium nitrate, ethylene and bromine interact to produce the compound Br·CH₂—CH₂·NO₂, of which no precursor such as Br·NO₃ is known.

Further theoretical evidence in support of the hypothesis of initial reaction of a positive halogen ion, and subsequent neutralization of the resultant charge by any approaching negative ion, may be drawn from the fact that in all these unsymmetrical addition-reactions the halogen atom combines with the ethylenic substances exclusively in the position required for the neutralization of the electronic charge on a negatively polarized carbon atom (see p. 178, also Chapter XI). For example, cinnamic acid will add on EtO·Cl, in 6% yield, by the action of dry chlorine in ice-cold ethyl alcoholic solution, giving only the α-chloro compound Ph·CH(OEt)·CHCl·COOH.¹

- (c) Hydrogenation. For the hydrogenation of olefines, Burton and Ingold have proposed a similar mechanism, involving consecutive addition of two protons and not simultaneous addition of two neutral atoms of hydrogen (compare p. 137). They imagine the reducing agent, such as a metal and an acid, to react in three stages:
- (i) A proton from the acid is added to an activated double bond to form a positively-charged ion, e.g.

$$\dot{H} + \bar{C}H_{\bullet} - \dot{C}H_{\bullet} \longrightarrow H \cdot CH_{\bullet} - \dot{C}H_{\bullet}.$$

¹ Jackson and Pasiut, J.A.C.S., 1927, **49**, 2071. Compare, however, Ogg, J.A.C.S., 1935, **57**, 2727.

¹ J. Chem. Soc., 1929, 2022.

(ii) This organic kation then acquires two electrons from the metal and is converted into an anion, e.g.

$$H \cdot CH_2 - \overset{+}{C}H_2 + Zn \longrightarrow Zn^{++} + H \cdot CH_2 - \overset{-}{C}H_2$$

(iii) The resulting anion, H·CH₂—CH₂, then adds on a second proton and becomes a neutral saturated compound, e.g.

$$H \cdot CH_2 - \overline{C}H_2 + H \longrightarrow H \cdot CH_2 - CH_2 \cdot H.$$

This mechanism can be applied to the reduction of other classes of organic compounds such as aldehydes and ketones, nitro- and nitroso-compounds, azo-compounds, and so forth. Even the complex reduction of a ketone to a pinacone can be formulated on similar lines as depending on the mutual neutralization of organic kations and anions, e.g.

$$(CH_3)_2\overset{+}{C}-O-H + (CH_3)_2\overset{-}{C}-O-H \longrightarrow (CH_3)_2C-O-H$$

The non-polar mechanism suggested in Chapter VII is, however, equally applicable to many of these reductions ¹; a decision as to the precise mechanism of any reduction process can therefore only be inferred (as in the case of halogen-addition) after a detailed examination has been made of the experimental conditions under which the reaction takes place.

The Sign of Ionization of Double Bonds.

Ionization of a double bond generally takes place in such a way that the more "negative" atom (i.e. the atom with the larger nuclear charge in a series such as C, N, O, F) secures the extra electron and thus acquires a negative charge. Thus a ketone is always ionized as follows:

$$CH_3$$
 $C=0 \Rightarrow CH_3$ $C+\bar{O}$

¹ Cf. the foothote in Burton and Ingold's paper.

The permanent dipole moment of the carbonyl group, which amounts to 2.7×10^{-18} e.s.u. even in the stable covalent form, is therefore increased in magnitude but not altered in sign, when the double bond is activated to a semi-polar form.

Addition-compounds are then formed according to the usual rule, by adding a positively-charged proton to the negatively-polarized oxygen and the "rest of the reagent" to the carbon. For the same reason, the ionization of double bonds between nitrogen and oxygen or carbon will proceed in opposite senses as follows:

$$-N=0 \Rightarrow -N-0$$

 $-N=0 \Rightarrow -N-0$

No similar general rule can be given for the ionization of unsaturated hydrocarbons, where the double bond unites two atoms with equal nuclear charges. It is therefore a matter of indifference which of the two carbon atoms of ethylene acquires a positive and which a negative charge; but, in the unsymmetrical derivatives of ethylene, a directive influence may be exercised by the substituents. Thus Markownikow's rule postulates that "the negative constituent attaches itself to the carbon atom which carries the smallest number of hydrogen atoms". The hydrogen of a reagent, HX, therefore tends to build up a complete methyl group, or failing that a methylene group, rather than itself take up a "tertiary position", e.g. iso-butylene and hydrogen bromide tend to give tert-butyl bromide rather than iso-butyl bromide.

$$CH_3$$
 $C=CH_2 + HBr \rightarrow CH_3$ $CBr-CH_3$, not CH_3 $CH-CH_2Br$

Markownikow's attempt to summarize the experimental evidence in one simple rule is subject to many exceptions. In particular, by-products are almost always formed in opposition to the rule, in proportions which depend upon the experimental conditions under which the reaction is carried out. Thus Michael in 1899 reported that propylene gave only about one-quarter per cent. of by-product with hydrogen iodide, but 25% with iodine chloride and 42% with iodine bromide.

```
Main Product

CH<sub>3</sub>·CHI·CH<sub>3</sub> (99½%) 

CH<sub>3</sub>·CH<sub>2</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH<sub>4</sub>·CH
```

Recently Kharasch, McNab and Mayo² have shown that the normal action between pure propylene and hydrogen bromide leads to a quantitative yield of *iso*-propyl bromide, CH₃·CHBr·CH₃, but, after exposure to the air, or in the presence of a peroxide such as benzoyl peroxide, *normal* propyl bromide, CH₃·CH₂·CH₂·Br, is formed instead. Solvents, and all other catalysts, apparently act indirectly, affecting the composition of the reaction-product only through their action on peroxides.

Michael considered the preferential formation of one isomeride in any unsymmetrical addition process to be due to an *unequal distribution of polar combining affinity*, but a more precise interpretation of the phenomenon can now be given in terms of the electronic theory of valency (see Chapter XI).

Addition of hydrogen bromide to acrylic acid, and most of its homologues 3 takes place quantitatively in a sense opposed to Markownikow's rule, but is again dependent upon experimental conditions. 4 There can be little doubt that this action depends, as Thiele suggested (p. 374), on the addition of hydrogen bromide to the ends of a conjugated chain, giving rise to the enolic form of

¹ J. Prakt. Chem., 1899, **60**, 286, 410; 1892, **46**, 205, 345, 452. ² J.A.C.S., 1933, **55**, 2531; cf. J.A.C.S., 1934, **56**, 1212.

³ Cf. Boorman, Linstead, and Rydon, J. Chem. Soc., 1933, 569. ⁴ Linstead and Rydon, J. Chem. Soc., 1934, 2001; Walker and Lumsden, J. Chem. Soc., 1901, 79, 1191.

180 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

the bromo-acid, which at once reverts to the normal form:

CH₂—CH—C
$$\stackrel{O}{\downarrow}$$
 + HBr \rightarrow Br·CH₂—CH—C $\stackrel{O}{\downarrow}$ OH \rightarrow Br·CH₃—CH

This topic will, however, be discussed more fully in a later chapter (Chapter XV).

Classification of the Reagents of Organic Chemistry.

In earlier paragraphs it has been shown that many organic reactions are brought about by ions, such as the CN ion of KCN, or the CH₃ residue which is left when I is eliminated from CH₃I or SO₄ from Me₂SO₄. The action of reagents such as these is obviously determined very largely by the sign of the ions to which they owe their activity. Lapworth 1 and Robinson 2 have, therefore, classified the common organic reagents as anionoid or kationoid in accordance with the following definition:

"If the formation of a covalent bond in the complex A-B takes place by the union of A with B then A is anionoid and B is kationoid."

The following table, due essentially to Robinson,² summarizes the ionic classification of the more important groups and reagents of organic chemistry. It includes, not only active ions of the types A and B such as CN and CH₂, but also covalent compounds, such as CH₃I and Cl₄ from which they may be derived, and (on the opposite side of the table) the covalent compounds with which they interact.

¹ Mem. Manchester Phil. Soc., 1925: Nature, 1925, 115, 625.

² See Solvay Reports, 1931, 434; cf. Ingold, Chemical Reviews, 1934, 15, 265-73.

Anionoid (Give up electrons)

(i) Active anions, e.g. OH, NH₂, CN, CH(COOEt)₂, etc.

- (ii) Reducing agents, including all metals and ions which can give up electrons, e.g.
 Na, Mg, Fe, Fe++ and
 Fe(CN)_e
 ':
- (iii) Alkyl and aryl residues of organo-metallic compounds, e.g. Et in Et·Mg·Br, and R·C—C in R·C—C·Na.
- (iv) Donor molecules containing lone pairs of electrons, e.g. NH₃₁, C₅H₅N, H₂O, Et₂O, CH₃·CO·CH₃ (on the oxygen) R·SH, R₂S, etc.
- (v) Unsaturated carbon of olefines and aromatic hydrocarbons, e.g. C₂H₄, C₆H₆.

Kationoid (Take electrons)

- (i) Protons and sources of protons, e.g. acids and pseudoacids.
 - Diazonium ions.1

Kations from pseudo-bases, e.g. cotarninium.

- (ii) Oxidizing agents, including the halogens, ozone, peroxides, and ions which can accept electrons; CrO₃-, Fe+++, Fe(CN)---, MnO₄-, etc.
- (iii) Alkyl residues from esters, e.g. Me in MeI or Me₂SO₄. Alkyl residues from quaternary ammonium salts and bases, e.g. Me from Ph·NMe₃OH, etc.
- (iv) Acceptor atoms, ions and molecules including metallic atoms and ions which are capable of co-ordination with H₂O, NH₃, etc., e.g. ZnCl₂, PtCl₄.
- (v) Atoms which carry, or easily acquire, positive charges, e.g. in semi-polar bonds, C in >CO, —COOEt, —CN; S in SO₃, H₂SO₄, NaHSO₃; N in —NO, —NO₂, HNO₃.

$C_6H_5 \cdot \stackrel{+}{N_9Cl} \longrightarrow C_6H_5 \cdot N:N \cdot OH$ $C_6H_5 \cdot N:N \cdot OH + H:NH \cdot C_6H_5 \longrightarrow C_6H_5 \cdot N:N \cdot NHC_6H_5.$

The condensation must be formulated as an ionic reaction, in which the diazo-compound gives up a hydroxyl-ion and leaves a positive residue, C₆H₅·N:N, whilst the aniline loses a proton and leaves a

¹ Diazonium salts are prepared in acid solutions, but are generally used in presence of sodium acetate. Under these conditions, the relatively-stable diazonium salts, which are not directly reactive, are converted into diazohydroxides, which undergo condensation with aniline, etc.

182 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

(a) Anions and Kations. In accordance with the preceding definition, anionic activity is determined by the readiness of an anion A to share its electron with another atom. This activity is at a minimum in the more stable anions, but increases progressively as the ion becomes less and less like an inert gas in the series

$$\bar{F}\!<\!\bar{O}H\!<\!\bar{N}H_{1}\!<\!\bar{C}H_{3}$$

Conversely, kationic activity is determined by the readiness of a kation B to share an electron-pair belonging to A. It is therefore least in the most stable kations and disappears completely in ions such as Na and K, which are almost incapable of forming covalent compounds, but is very intense in unstable residues such as CH₃.

The stability of an anion can be measured quantitatively by means of its *electron-affinity*, as follows:

Cl +
$$e = \overline{Cl}$$
 + 88,000 cal. = 3.8 volts
Br + $e = \overline{Br}$ + 80,000 cal. = 3.5 volts
I + $e = \overline{I}$ + 71,000 cal. = 3.1 volts

Conversely, the stability of a kation can be expressed quantitatively by recording the *ionization-potential* of the metal as follows:

H
$$\rightarrow$$
 H + e - 312,000 cal. \equiv - 13.54 volts
Li \rightarrow Li + e - 125,000 cal. \equiv - 5.4 volts
Na \rightarrow Na + e - 119,000 cal. \equiv - 5.2 volts
K \rightarrow K + e - 100,000 cal. \equiv - 4.3 volts

negative residue, C_0H_5 : these very unstable residues then unite to form the diazo-amino-compound.

Similar considerations apply to the aliphatic diazo-compounds,* which can only react in the form which contains tervalent nitrogen, although this may be able to break down either into ions or into neutral free radicals.

^{*} Bradley and Robinson, J. Chem. Soc., 1928, 1312.

These quantities refer, however, to the ionization of gaseous atoms and are only related indirectly to the conversion of liquid or solid elements into hydrated ions in aqueous solutions. For instance, the ionization of molecular chlorine, bromine and iodine depends also on the energy of dissociation of the molecules into atoms as follows:

Cl₃ = 2Cl - 106,000 cal.
$$\therefore$$
 Cl₃ + 2 e = 2Cl + 70,000 cal.
Br₃ = 2Br - 46,000 cal. \therefore Br₃ + 2 e = 2Br + 114,000 cal.

$$I_2 = 2I - 36,000 \text{ cal.}$$
 $\therefore I_2 + 2e = 2I + 106,000 \text{ cal.}$

Even these figures, however, are misleading if applied to the mutual displacement of halogens from their salts, since they refer to the production of gaseous ions from gaseous molecules and are not therefore applicable directly to the liquid or solid halogens and to the hydrated halide ions in aqueous salt solutions.

(b) Oxidizing and Reducing Agents. In general, reducing agents give up electrons and oxidizing agents accept them. Thus the oxidation of a ferrous salt and of a ferrocyanide is represented by the equations:

$$Fe \longrightarrow Fe + e : FeC_{\bullet}N_{\bullet} \longrightarrow FeC_{\bullet}N_{\bullet} \longrightarrow + e$$

In the same way the oxidation of hydroquinone to quinone in alkaline solutions can be expressed by the equation

$$O \cdot C_0 H_4 \cdot O \longrightarrow O : C_0 H_4 : O + 2e$$
.

All atoms, molecules or ions which can act as reducing agents are therefore classed as anionoid reagents in virtue of their ability to act as *electron-donors*, even if the electrons are transferred to other atoms and not merely shared. Conversely, all oxidizing agents are classified as kationoid reagents in virtue of their ability to act as *electron-acceptors*, whether they act by electron-sharing or by electron transfer.

When used as oxidizing agents the halogens absorb two electrons, $Cl_1 + 2e = 2Cl$, and therefore are included amongst the kationoid reagents as electron-acceptors. When used as halogenating agents they usually yield one ion

and one covalent molecule; but their classification as kationoid reagents is justified by the fact that initial ionization in the sense of the equation $Cl_2 \rightleftharpoons Cl + Cl$ gives rise to one active kationoid ion and one inactive ion. The formation of one active ion from each molecule of halogen is in contrast to its photochemical activation which gives

rise to one active kationoid ion and one inactive ion. The formation of one active ion from each molecule of halogen is in contrast to its photochemical activation which gives rise to two active atoms. This produces a difference in the order of reaction, which makes it possible to discriminate between the two types of reaction. Thus, in polar reactions we have:

 $\mathrm{Br_2} \rightleftharpoons \overset{+}{\mathrm{Br}} + \overset{-}{\mathrm{Br}}$, whence $dk/dt = k_3[\overset{+}{\mathrm{Br}}] \approx k_4[\mathrm{Br_2}]$, but in non-polar reactions

 $Br_2 \rightleftharpoons 2Br$, whence $dk/dt = k_1[Br] = k_2[Br_2]i$.

In the former case the velocity is directly proportional to the concentration of the halogen, but in the latter it is proportional to the square root of this concentration.

- (c) Alkylating Agents. Alkylation by means of an organometallic compound leads to the formation of metallic kations and is therefore attributed to the formation of an organic anion, such as $\overline{CH_3}$, which acts as an anionoid reagent; but alkylation with an alkyl iodide or sulphate leads to the formation of an inorganic anion and an unstable organic
- kation, such as CH₃, which acts as a kationoid reagent.
- (d) Donor and Acceptor Molecules. In the formation of co-ordination-compounds and similar reactions, the donor-molecule of NH_3 , H_2O , etc., is classed as an anionoid reagent and the acceptor molecule or ion of Co, Pt^{++} , etc., is classed as a kationoid reagent.
- (e) Unsaturated compounds, which act by forming semi-polar bonds, are classed as kationoid if the positive pole is more reactive than the negative pole, but as anionoid if the negative pole is more reactive. Thus the carbonyl compounds are strongly kationoid, since they are ionized invariably in the sense of the scheme

 $>C=O< \Rightarrow >C^{\dagger}-O$, in which the negatively charged

oxygen ion is relatively stable (as in the ions OH and OEt), whereas the positively charged carbon is very unstable and therefore intensely reactive.

On the other hand, the olefines when ionized yield two very reactive poles $>C==C< \Rightarrow >C-$ Of these, the negative is the more reactive, as in the triphenylmethyl series (p. 124) where the anion of Ph_3C Na is more reactive than the kation of Ph_3C Cl. The olefines are therefore classed as anionoid substances, and this classification is justified by the experimental fact that they interact more readily with kationoid reagents, such as chlorine or nitrosyl chloride, than with anionoid reagents, such as sodamide.

The ease with which addition to olefines can usually be brought about thus depends upon the high reactivity of the polar centres in the activated form; and the differences between the characteristic reactions of the two principal types of unsaturated compounds, although wide enough to produce a different behaviour towards every reagent except hydrogen, can be interpreted immediately in terms of the electronic configurations of the atoms forming the double bond, and of the unequal affinities which these atoms acquire as a result of the polar activation of this bond. Indeed, in all addition-reactions of unsaturated compounds, the characteristics displayed are those of the semi-polar, activated, structure and not of the normal wholly-covalent form.

The Amphoteric Character of Carbon.

Carbon occupies a central position in the periodic classification of the elements and yields neutral compounds (CH₄, CCl₄, etc.) both with hydrogen and with chlorine. These form stable covalent molecules, which contain complete octets of electrons, and have no appreciable tendency to break down into ions. In organic reagents, however, carbon is associated with radicals (like the magnesium in the Grignard reagent, and the sulphate radical in dimethyl

sulphate), which are liberated as ions, so that it must itself form part of an organic anion or kation, at least momentarily during the course of chemical reactions.

The positive hydrocarbon ions, derived from compounds, which contain a group that has a tendency to break away in the form of a stable anion (e.g. I⁻ from CH₃I), interact with atoms such as oxygen which readily acquire a negative charge: they are therefore used in the alkylation or arylation of the sugars and oximes, and for the preparation of ethers and esters by interaction with the metallic derivatives of alcohols and acids.

The negative hydrocarbon ions, which are derived from organo-metallic compounds from which metallic kations tend to separate, are used as reagents to attach alkyl or aryl radicals to positively polarized atoms of carbon, as in the preparation of tertiary alcohols from ketones or esters by means of the Grignard reagent, or in syntheses with malonic ester or acetoacetic ester.

Like the free neutral radicals, free ions of carbon compounds are all highly reactive, and tend to combine with other ions to form stable wholly-covalent molecules. As a general rule it can be stated that the negative alkyl ions are even less stable and more reactive than the positive alkyl ions, since the reagents which produce them are less numerous and less easy to prepare. Experimental evidence, which will be detailed below, shows, however, that both positively and negatively ionized hydrocarbon radicals are capable of independent existence for periods of time which can vary from the almost negligible instants of actual chemical change to periods of almost indefinite duration.

Organic Kations.

(a). The Carbonium Salts.

The simple alkyl halides are non-electrolytes, and yield alkyl ions only as unstable intermediate-products of chemical change. When, however, the three hydrogen atoms of methyl iodide are replaced by phenyl-groups, the triphenyl-methyl radical and the halogen separate so easily that the

halogen-derivatives possess all the essential characteristics of salts. Thus triphenylmethyl bromide, when dissolved in liquid sulphur dioxide at 0° , is a better conductor than potassium iodide in the same solvent 1; and, at dilutions below N/500, it is actually a better conductor than potassium iodide is in aqueous solutions at 18° C. This remarkable fact is clearly established by the following comparison:

The chloride as well as the bromide is an excellent conductor when dissolved in liquid hydrogen cyanide,² and the bromide is also a fair conductor when dissolved in pyridine.³

CPh₃Cl in HCN. .
$$V = 121$$
 393 1,668 $\lambda = 45$ 75 115

CPh₃Br in HCN . $V = 130$ 406 1,790 $\lambda = 166$ 202 211

CPh₃Br in pyridine . $V = 320$ 520 1,400 3,175 $\lambda = 18.9$ 27.7 29.7 44.8

By introducing the anisyl, naphthyl or diphenyl radicals in place of phenyl, the coefficient of ionization is still further increased so that even the chlorides become good electrolytes. The degree of dissociation or "coefficient of ionization", α , for some of these compounds at V = 100 has been calculated as follows.4

¹ Walden, Ber., 1902, 35, 2023.

² Gomberg and Sullivan, J.A.C.S., 1922, 44, 1820.

⁸ Hantzsch and Meyer, Ber., 1910, 43, 337.

⁴ Gomberg and Sullivan, J.A.C.S., 1922, ⁴⁴, 1818; Gomberg and Nishida, J.A.C.S., 1923, ⁴⁵, 206.

			Degree of D	
Kation.			Chloride.	Bromide.
CPh_3^+			0.13	0.61
β -C ₁₀ H ₂ ·CPh ₂ +.			o.18	0∙58
α -C ₁₀ H ₇ ·CPh ₂ +.			0.30	o∙58
o-MeO·C ₆ H ₄ ·CPh ₂ +			0.47	0.57
$(C_{\bullet}H_{\bullet}\cdot C_{\bullet}H_{\bullet})_{3}C+$			o·45	
cf. $N(CH_3)_4$ + .				o·56

It will be seen that the degree of dissociation of the chlorides is less than that of the bromides, but that the latter are dissociated to a slightly greater extent than tetramethylammonium bromide. It is also noteworthy that the molecular conductivity of α -naphthyldiphenylmethyl bromide in HCN is comparable with that of an aqueous solution of potassium hydroxide, or about *twice* the conductivity of the alkali halides in water.

Von Baeyer proposed to call these compounds carbonium salts, and introduced a special symbol for the carbonium bonds as follows, Ph₂C~~~O·SO₂·OH. The term "carbonium" appears to be fully justified by the analogy between the behaviour of these carbon compounds and that of the alkylated ammonium salts, but the zigzag symbol is no longer needed, since the ionization of these salts can be represented by the ordinary symbols as follows:

 $Ph_3C \cdot Br \rightleftharpoons Ph_3C + Br$. Whereas, however, many metallic salts are completely ionized, the ionization of the carbonium salts is obviously reversible, like that of the acids and of the mercuric salts, since no valency-rule is violated when the ions unite to form a covalent compound of quadrivalent carbon.

(b) Chemical Properties of Carbonium Salts.

In general, alcohols are less efficient as bases than as acids. This is emphatically true of the phenols

$$\stackrel{-}{HO} + \stackrel{+}{C_0}H_0 \iff \stackrel{-}{C_0}H_0 O \cdot H \implies \stackrel{-}{C_0}H_0 O + \stackrel{+}{H};$$
Basic dissociation.

which are often described as acids (e.g. picric acid, carbolic acid, etc.). On the other hand, the alcohols of the triphenyl-

carbinol series are definitely basic in character. Thus they can be converted into the corresponding chlorides by merely dissolving in benzene and passing in hydrogen chloride, or by adding acetyl chloride to the solution.

$$CPh_3 \cdot OH + HCl = Ph_3C - Cl + H_2O$$

 $Ph_3C \cdot OH + Cl \cdot CO \cdot CH_3 = Ph_3C \cdot Cl + HO \cdot CO \cdot CH_3$.

The former process is just like a simple neutralization. The latter process would be expected to give rise to an ester, Ph₃C·O·CO·CH₃, but the "carbonium" radical has such a strong tendency to ionization that even the "ester" behaves as a salt and (as might be expected) the acetate is decomposed by hydrochloric acid giving rise to the chloride and free acetic acid.

(c) Solvation of Carbonium Salts.

The carbonium kation contains only a sextet of electrons, like the positive halogen-ions discussed in a previous paragraph (p. 184). It would therefore be expected to be extremely unstable. Nevertheless, the conductivity figures cited above show that the triphenylmethyl kation can persist in considerable concentrations when dissolved in a suitable ionizing solvent. Since these solutions show a characteristic absorption spectrum, which is not found in solutions in non-ionizing solvents, it has been suggested that the ion has been stabilized by solvation. Sidgwick has suggested that the kation in liquid sulphur dioxide, for example, might be formulated as (Ph₃C, SO₂) and

represented by an electronic formula such as Ph₃C:O:S:O:, where every atom has its complete octet of electrons, as

in sulphur dioxide, :0:S:O:.

$$\overline{O}$$
 $\stackrel{++}{S}$ $\stackrel{-}{O}$ $\stackrel{+}{C}$ $\stackrel{+}{C}$ $\stackrel{+}{C}$ $\stackrel{+}{O}$ $\stackrel{-}{S}$ $\stackrel{+}{O}$ $\stackrel{-}{C}$ $\stackrel{+}{C}$ $\stackrel{+}{C}$ $\stackrel{-}{C}$ $\stackrel{+}{C}$ $\stackrel{+}{C}$ $\stackrel{-}{C}$ $\stackrel{+}{C}$ $\stackrel{-}{C}$ $\stackrel{-}{C}$ $\stackrel{+}{C}$ $\stackrel{-}{C}$ $\stackrel{-}{C$

¹ Wieland and K. H. Meyer, Ber., 1911, 44, 2557.

² Walden, 3rd Solvay Report, 1928, p. 510.

(d) The Colour of Aromatic Kations.

"Free radicals" generally give rise to coloured solutions. Since, then, the colour and degree of dissociation of carbonium radicals both increase with dilution, it has been suggested that the development of colour is due to the dissociation of a colourless molecule into a coloured radical. This simple explanation is, however, untenable since (i) there is no quantitative relation between colour and degree of dissociation, e.g. of \(\beta\)-naphthyl diphenylmethyl in benzene, nitrobenzene and cyclohexane, (ii) the absorption spectrum of triphenylmethyl in benzene is different from that in sulphur dioxide, and (iii) the absorption spectrum of the triphenylmethyl radical is different from that of the triphenylmethyl kation produced by dissolving the carbinol in sulphuric acid. It has therefore been suggested that the monomolecular radical exists in two forms, one of which may be regarded as "quinonoid".

This view is extremely plausible, since the existence of two oppositely-charged ions has already been postulated. It would, then, be reasonable to suppose that the negative ion, which carries a full complement of electrons, may retain its original structure, but with a "lone pair" of electrons as a "scar" to mark the point from which the other ion has been detached. In the case of the positive ion, however, there is a shortage of electrons which obviously may be transferred to the part of the molecule which is most ready to tolerate it. Thus, it is most unlikely that the oxygen atom of dehydromesitol would be content with a sextet of electrons; we should, therefore, expect that the deficit of electrons in the positive ion would be transferred from oxygen to carbon, as in the "quinonoid" formula (II).

Structural changes of the above type are widely used to explain colour-changes on the formation of salts (i.e. of ions) from aromatic pseudo- or carbinol-bases, e.g. in the pyrane ring of the anthocyanidines, which comprise the flower pigments, and in the benzene ring of the triphenylmethane, acridine, and phenazine dyestuffs such as magenta, chrysaniline and safranine. In the case of triphenylmethyl itself, experimental evidence ¹ points to the occurrence of molecular rearrangement subsequent to the initial dissociation into free ions.

(Unstable kation.) (Stable anion.)

Ph₃C—CPh₃
$$\rightleftharpoons$$
 Ph₃C+ + Ph₅C-
(Colourless.)

H H

Ph

C=C
H

H H

(Yellow stable kation.)

Definite chemical evidence for the existence of quinonoid systems has been supplied by experiments which have demonstrated the mobility of the halogens in p-halogenated derivatives of triphenylmethyl. For example, tri-p-bromophenylmethyl chloride after solution in sulphur dioxide yields some chloro-dibromo-triphenylmethyl chloride 2 (BrC₆H₄)₃CCl \rightleftharpoons (BrC₆H₄)₂CBr (C₆H₄Cl). Again, the free radical, mono-p-bromotriphenylmethyl, contains an active atom of bromine, which can be removed by means of metallic silver 3 ; this can be explained most easily by means of the isomerization

The triphenylmethyl kations have one electron less than the free neutral radicals shown above; but there is no

¹ Gomberg and Sullivan, J.A.C.S., 1922, 44, 1812.

² Gomberg and Cone, Annalen, 1910, 376, 183.

³ Gomberg, Chemical Reviews, 1924, 1, 124.

a priori reason why electron movements within molecules should necessarily involve one or two electrons only.1

Organic Anions.

The spontaneously inflammable zinc alkyls have already been cited as compounds which may contain an alkyl anion. They represent, however, but one group out of a whole class of organo-metallic derivatives of varying degrees of stability, ranging from the moderately stable metallic acetylides, $R-C \equiv C \cdot Na$, like sodium acetylide to the reactive sodium triphenylmethyl, or the still more reactive compounds, sodium benzyl, $Na \cdot CH_2 \cdot C_6H_6$, and sodium ethyl, $Na \cdot C_2H_6$. These have been prepared by Schlenk,² by warming the alkali metal with the requisite mercury alkyl in an atmosphere of an inert gas such as nitrogen, and then distilling off the mercury under diminished pressure, e.g.

$$2Na + Hg(C_2H_5)_2 \rightarrow Hg + 2Na \cdot C_2H_5$$

but are also obtainable by the action of the metal directly on alkyl and aryl halides in ether or benzene.³

They are white crystalline solids, closely analogous to the metallic hydrides such as LiH. They decompose water, are spontaneously inflammable, and often react with air with explosive violence. The corresponding metallic aryls,

e.g. Ph₃C Na, are more stable and are deeply coloured, ranging in tint from deep red to violet or blue.

As long ago as 1903 it was suggested that these organometallic compounds were polar in nature. Their ionic structure has been confirmed by measurements of conductivity of solutions in ether, pyridine, and liquid

¹ Cf. Ingold, Trans. Faraday Soc., 1934, 30, 52.

² Cf. Schlenk's article in Houben-Weyl's "Methoden der Organischen Chemie".

³ Ziegler and Colonius, *Annalen*, 1930, **479**, 135; Gilman, Zoellner and Selby, *J.A.C.S.*, 1932, **54**, 1957.

⁴ Schlenk and Marcus, Ber., 1914, 47, 1664.

⁵ Schlenk and Holtz, Ber., 1916, **49**, 606; 1917, **50**, 276; Ziegler and Wollschnitt, Annalen, 1930, **479**, 123.

ammonia. Thus the equivalent conductivity of sodium triphenylmethyl in 0.05 N solution in liquid ammonia is 64. Triphenylmethane also behaves as a weak acid in liquid ammonia, since it is possible to prepare metallic aryls in the manner indicated by the following equation:

$$Ph_3C \cdot H + Na \cdot NH_3 \rightleftharpoons Ph_3C \cdot Na + NH_3$$
.

The system is, however, reversible and the equilibrium favours the free hydrocarbon.

The anionic character of triphenylmethyl in its metallic derivatives has been confirmed by the preparation of a curious salt, $(CH_3)_4NCPh_3$, in which five hydrocarbon residues are associated with one nitrogen atom. This organic salt is necessarily a strong electrolyte, completely ionized in all solvents. On the other hand, triphenylmethyl itself is ionized only reversibly in ionizing solvents,

 Ph_3C — $CPh_3 \rightleftharpoons Ph_3C^+ + CPh_3$. Thus, in liquid sulphur dioxide at o° the conductivities observed at varying dilutions were as follows:

$$V = 24$$
 43 66 98 137 165 230 461 2,138 $\lambda = 8.24$ 11.97 17.14 22.15 26.79 28.89 31.27 32.50 39.59

General Conclusions.

In concluding this chapter one must stress the fact that the great majority of organic ions are, like the neutral organic radicals discussed in Chapter VII, highly reactive entities, which form only under favourable conditions during the course of a chemical change. It has already been indicated (p. 94) that it is more difficult to dissociate a molecule into charged ions than into neutral radicals, but that ionization can be greatly facilitated by the use of a suitable solvent. This is usually a substance of high dielectric constant, in which the forces of electrical attraction between ions of opposite sign will be greatly dimin-

¹ Kraus and Rosen, J.A.C.S., 1925, **47**, 2739; Kraus and Kawamura, J.A.C.S., 1923, **45**, 2756; Schlenk and Ochs, Ber., 1916, **49**, 614.

194 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

ished. Even more important, however, is the anionoid character of the solvent, for substances such as water, ether, acetone and liquid ammonia, which can form co-ordination-compounds (solvates) with highly reactive kations, e.g.

with H, Co, facilitate ionization to a remarkable degree. It is by a similar action no doubt that traces of substances such as water can initiate an ionic reaction between two covalent molecules (p. 173).

It has already been mentioned (p. 83) that chemical reaction between two covalent molecules does not invariably occur upon collision of reactants possessing sufficient energy of activation. Other conditions must be satisfied also, and, for the reactions dealt with in this chapter, one may suggest that there is required the presence of a suitable catalyst, or solvent, by means of which the electrical circuit can be completed in and around the reacting molecules.

CHAPTER IX

ACIDITY

The Nature of Acidity.1

One of the most striking properties in the whole range of chemistry and indeed one of the first chemical properties to be noticed after combustion is that of acidity. Its origin was not clear until Davy, in 1816, proved the absence of oxygen in hydrochloric acid, and thus overthrew Lavoisier's theory that oxygen was the only "acid-producer" and was therefore an essential element in all acids. In its place, Davy advanced the view that hydrogen in an acid might be considered as acting the part played by a metal in the salts derived from it.²

The idea that acids are salts of hydrogen was stated in a definite form in Laurent's Chemical Method, published after his death in 1853.3 It received a new interpretation with the coming of the theory of electrolytic dissociation, since it then appeared that acids are characterized by the production of hydrogen ions in solution, just as sodium salts give rise to sodium ions. Acidity is therefore commonly expressed in terms of hydrogen ion concentration, but in the peculiar form of an index which grows less as the acidity increases, since the factor used is the symbol

$$p_{\rm H} = -\log_{10}[\rm H^+]$$

This factor is determined most accurately in an aqueous solution by measuring the hydrogen ion potential, with the help

¹ Cf. Walden, "Salts, Acids and Bases", G. F. Baker Lectures, Cornell University, 1929.

² Works, V, 514.

[&]quot; Chemical Method", Cavendish Society Reprints, 1854, p. 41.

of a standard hydrogen-gas, quinhydrone or glass electrode. The p_H of the solution is therefore defined more precisely by the Nernst electrode equation, viz.

$$E = \frac{RT}{F} \times p_{H} = \frac{RT}{F} \log \frac{1}{\alpha_{H}}$$

where E, the electrode potential measured in volts, is referred to the normal hydrogen electrode as the standard zero, and the hydrogen ion concentration, [H+], has been replaced by the hydrogen ion activity, $\alpha_{\rm H}+$. This activity is related to the concentration by a factor, γ , which is termed the activity coefficient, so that $\alpha_{\rm H}=\gamma[{\rm H}^+]$. Since this factor is very nearly equal to unity for ions and molecules at low concentrations in water, no appreciable error is introduced by referring acidity to hydrogen ion concentration in weak organic acids.

An Extended Definition of Acids and Bases.

An acid may be defined, in terms of modern theories of atomic structure, as a *proton-donor*. Thus it is characteristic of an acid to yield hydrogen ions in solution, or, more strictly, to give a proton to a molecule of water and thus yield an oxonium ion

$$OH_1 + HCl = OH_1 + Cl.$$

Similarly, on electrolysis an acid liberates positively-charged hydrogen atoms at the kathode, $HCl \rightarrow H + Cl$, where the protons acquire electrons from the electrode and are converted into hydrogen gas,

$$_{2H}^{+} + _{2e} = H_{\bullet}$$

The same change takes place when an acid dissolves a metal, since in this process the acid is electrolysed and the resulting protons are discharged by electrons from the metal

$$HCl \rightarrow \overset{+}{H} + \overset{-}{Cl}; \ 2\overset{+}{H} + Zn \rightarrow \overset{+}{Zn} + H_1.$$

In order that an acid may be isolated as a chemical compound, the proton-donor must be an electrically-neutral molecule; but the same property is also found in

¹ Cf. W. M. Clark, "The Determination of Hydrogen Ions", H. T. S. Britton, "Hydrogen Ions", etc.

² Cf. Lewis and Randall, "Thermodynamics", pp. 264, 332, etc.

(i) Negatively-charged ions of a polybasic acid, e.g.

$$HSO_4 \rightleftharpoons H + SO_4$$

(ii) Positively-charged ions of "onium" salts, formed by the reversible addition of a proton to a base

$$\overset{+}{O}H_3 \rightleftharpoons \overset{+}{H} + OH_4.$$

The converse property of basicity, which was formerly measured in terms of hydroxyl-ion concentration [OH⁻], is now defined ¹ as ability to take up a proton. The class of proton-acceptors includes:

(i) Electrically-neutral molecules of water and of nitrogenous bases, which are converted into positively-charged ions by taking up a proton, e.g.

$$NH_3^+ + H \rightleftharpoons NH_4^+$$

(ii) Negatively-charged ions, which can discharge a proton, the most important of these being the hydroxyl ion,

$$OH + H \rightleftharpoons OH_{\bullet}$$

together with compounds containing such ions, of which the most important are the alkalis, e.g.

$$\overset{+}{K}\overset{-}{OH} + \overset{+}{H} \rightleftharpoons OH, + \overset{+}{K}$$

(iii) Positively-charged ions of diacid-bases, such as hydrazine, which have already accepted one proton, e.g.

$$\overset{+}{N}H_{1}\cdot NH_{1} + \overset{+}{H} \rightleftharpoons \overset{+}{N}H_{3}\cdot \overset{+}{N}H_{3}.$$

In general, both acids and bases can be defined through the general equation:

$$Base + Proton = Acid$$
or $B + H^+ = A$

where B is described as the conjugate base to the acid A.2

The free proton, however, resembles an α -particle in that it consists of a naked nucleus with no planetary electrons; it is therefore very unstable and always attaches itself to the

¹ Lowry, Chemistry and Industry, 1923, **42**, 43 Trans.; Faraday Soc., 1930, **26**, 47; cf. Brönsted, Rec. Trav. Chim., 1923, **42**, 718; Chemical Reviews, 1928, **5**, 231-338.

² Wynne-Jones, Proc. Roy. Soc., 1933, A140, 440.

electronic shell of another atom or molecule, e.g. to a molecule of water to form an oxonium ion, $\overset{+}{H} + OH_2 \rightarrow \overset{+}{O}H_3$. Fajans, for example, has calculated that the heat of hydration of free protons is 232,000 cals. per gram-atom, corresponding with a molar concentration of 10^{-150} of free H^+ ions in an aqueous solution! The ionization of an acid is, therefore, not a mere emission of a proton, but can only take place when there is a base to receive it, e.g.

$$OH_3 + HCl \rightleftharpoons OH_3 + Cl$$
; not $HCl \rightleftharpoons H + Cl$.

It is, therefore, indistinguishable from the neutralization of an acid,² which also consists in a transfer of a proton from one base to another.

Both processes can then be represented by a general scheme, as follows:

$$acid_1 + base_2 \rightleftharpoons acid_2 + base_1.*$$

Thus the neutralization of acetic acid by aniline, which occurs to the extent of about 50% only, is represented by the balanced equation:

$$CH_3 \cdot CO \cdot OH + C_6H_6 \cdot NH_2 \rightleftharpoons CH_3 \cdot CO \cdot O + C_6H_6 \cdot NH_3$$

Since the action is reversible it is clear that the acetate ion can accept a proton, and is therefore a weak base, whilst the anilinium ion can give a proton and is therefore a weak acid. Moreover, since the reversible action is so evenly balanced, it is clear that the acid and basic strength of the ions on the right-hand side of the equation must be about equal to that of the molecules on the left-hand side. This is, however, no longer true when stronger acids and bases are used, as in the equation

$$HCl + NH_3 \rightleftharpoons NH_4 + \overline{Cl}$$
.

The ammonium ion in this equation still possesses weak acid properties, and acts as a catalyst for prototropic change (p. 313); but the chloride ion is almost entirely devoid of

¹ Ber., 1919, 21, 549, 709.

² Cf. Hantzsch, Ber., 1927, 60, 1933.

^{*} Brönsted, Chemical Reviews, 1928, 5, 231-338.

basic properties, since the action $Cl + OH_3 \rightarrow HCl + OH_3$ is almost non-existent. The chloride ion is, therefore, a much weaker base than water, and has no appreciable catalytic activity in prototropic change (p. 313).

Effect of the Solvent.

The definitions given above are not confined to aqueous solutions and therefore assign no special function to the hydroxyl ion. An exchange of protons may indeed be carried out in any solvent; but, as Wynne-Jones has pointed out, the relative strengths of two acids may be reversed in two different solvents.

Thus in water o-nitrobenzoic acid has a dissociation constant of 6.2×10^{-3} and 3:5-dinitrobenzoic acid a dissociation constant of 1.6×10^{-3} , whilst in ethyl alcohol the respective constants are 2.42×10^{-9} and 8.16×10^{-9} .

Since the ionization of any acid A in a solvent S is due to a reaction of the type $A + S \rightleftharpoons SH + B$, it would appear at first sight that a change of solvent should affect all acids similarly. There are, however, two factors that complicate the relations:

- (i) the electrostatic action between **SH** and **B** will depend upon the solvent;
- (ii) the chemical potentials of A and B may vary differently with change of S.

The second factor is difficult to estimate, but it has been suggested by Bjerrum and Larsson that since B differs from A only by a proton, the variations in chemical potentials may be considered to be the same.

The first factor can be estimated by comparing the electrostatic potentials of the different ions in the two solvents.

Since work
$$= \frac{\varepsilon^2}{2r} \cdot \left(\frac{1}{D_1} - \frac{1}{D_2}\right)^*$$

must be done to transfer an ion of charge ε and radius τ

¹ Proc. Roy. Soc., 1933, A 140, 440.

² Zeit. phys. Chem., 1927, 127, 358.

^{*} Born, Zeit. für Physik, 1920, 1, 45.

from a solvent of dielectric constant D_1 , to one of dielectric constant D_2 , Wynne-Jones deduced that, with change of solvent, the relative change of dissociation constant (K_r) of an acid compared with a standard acid, should be expressed by the equation

$$\Delta \log K_r = \frac{\varepsilon^2}{2kT} \left(\frac{2z+1}{r} - \frac{2z_0+1}{r_0} \right) \cdot \Delta \left(\frac{1}{D} \right)$$

where z_{ε} and $z_{0\varepsilon}$ are the charges carried by the conjugate bases of the acid considered, and the standard acid, and r and r_{0} are their respective radii.

This deduction was verified by demonstrating that a linear relationship existed between $\log K_r$ and $\frac{\mathbf{I}}{D}$ for a wide range of different acids.

It is therefore possible to split up the relative strength of an acid into two terms:

- (i) that due to electrical forces, and (ii) the intrinsic strength; the latter being equal to the relative strength in a hypothetical solvent for which $\frac{\mathbf{I}}{D} = 0$.
- G. N. Lewis has given a still wider definition of acidity than that of Lowry and Brönsted, for he has stated 2 that:

"A basic substance is one which has a lone pair of electrons which may be used to complete a stable group of another atom . . . an acid substance is one which can employ a lone pair from another molecule."

According to this definition, any tendency to form an "onium" salt (e.g. oxonium, ammonium, sulphonium) is an indication of basicity, even if it is manifested only in reactions which do not involve hydrogen ions, such as

$$N(CH_3)_3 + CH_3 \stackrel{+}{I} = \stackrel{+}{N}(CH_3)_4 \stackrel{-}{I}$$

 $NH_3 + BF_3 \longrightarrow \stackrel{+}{N}H_3 - BF_3$

or

On the whole, however, it is probably better to limit the

¹ Proc. Roy. Soc., 1933, A 140, 443.

^{2&}quot; Valence", 1923, p. 142.

terms acidic and basic to donors and acceptors of protons and not to extend them to all acceptors and donors of electrons.

Origin of Acidity and Basicity.

- (a) Valency. The basicity of a molecule or ion is defined by its ability to accept a proton from an acid. This property depends directly on the presence of "lone pairs" of electrons which can be shared with a proton. Thus G. N. Lewis points out that:
 - "With the exception of the elements of the argon type, every atom which possesses one or more lone pairs of electrons shows a tendency to increase its valence by adding hydrogen ion or some corresponding radical . . . Thus, starting with the oxide ion, the addition of one hydrogen ion produces the ion $\stackrel{-}{\text{OH}}$, the addition of two produces water, and the addition of three produces the oxonium ion $\stackrel{+}{\text{OH}_3}$."
- (i) Successive protons are added with increasing reluctance, as the negative charge on the ion decreases or the positive charge increases. Basicity therefore decreases progressively in a series such as

$$0 \longrightarrow 0H \longrightarrow 0H_2 \longrightarrow 0H_3 \longrightarrow 0H_4$$

just as acidity decreases in series such as

$$H_{\bullet}SO_{\bullet} \rightarrow HSO_{\bullet} \rightarrow SO_{\bullet}$$

or

$$H_3PO_4 \rightarrow H_2PO_4 \rightarrow HPO_4 \rightarrow PO_4$$

When lone pairs of electrons are no longer available the basicity of a molecule or ion disappears abruptly. The quadricovalent hydrides

are therefore no longer basic, although they exhibit increasing acidity as the positive charge increases. Thus the ammonium ion has already been classified as a weak acid (p. 198) and the ion, OH_4 , must be an even stronger acid than the ion, OH_3 , to which the acidity of aqueous acids is mainly due.

202 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

(ii) Basicity depends, however, not only on the presence of lone pairs of electrons, but on their ability to accept a proton. This property is by no means inherent, since it is not possessed by the lone pairs of the inert gases, where the stability of the completed octet causes an abrupt disappearance of basic properties in the final member of the series:

$$[C] \longrightarrow [N] \longrightarrow O \longrightarrow F \longrightarrow Ne.$$

Ability to accept a proton therefore decreases as the number of lone pairs increases, as in the series:

$$NH_1 \rightarrow OH \rightarrow F$$

where the effect of unequal charges on the atom has been eliminated by selecting a series of univalent anions for comparison. Thus \overline{F} is the ion of a strong acid and has therefore only very weak basic properties; \overline{OH} is a strong base, as in the alkalis; and \overline{NH}_2 is so strongly basic that it cannot exist in presence of water, where it decomposes at once according to the equation $\overline{NH}_2 + \overline{OH}_2 \longrightarrow \overline{NH}_3 + \overline{OH}$.

The presence or absence of lone pairs of electrons and their loss of "sharing-power" as they approach a complete octet therefore afford a simple explanation of

- (i) The strong acidity of the univalent hydrides, FH, ClH, BrH, IH.
- (ii) The weakly acidic or amphoteric character of the bivalent hydrides, OH₂, SH₂, TeH₃.
- (iii) The mainly basic character of the tervalent hydrides, NH₂, PH₂.
- (iv) The complete neutrality of the quadrivalent hydrides, CH₄, SiH₄.
- (b) Polarizability. The properties of acidity and basicity can also be linked up with computations of ionic and molecular polarizabilities by Fajans, ¹ C. P. Smyth ² and

¹ Zeit. für Physik, 1924, 23, 1; Trans. Faraday Soc., 1927, 23, 357, 375, 408.

² Phil. Mag., 1924, 47, 530; 1925, 50, 361.

others. High polarizability, which is disclosed by a high molecular refraction, as in NH₂, is an indication of large atomic dimensions. More exactly, it indicates that the valency electrons move in extensive orbits and are con-

trolled by weak restraining forces; they therefore occupy a large sphere of influence, and consequently their co-ordinating, or basic, powers will be great. Any increase of nuclear positive charge will restrain the valency electrons and so decrease their radius of action. When the effect of the nuclear positive charge becomes great, as in HCl, any further co-ordination of protons is prevented by the mutual repulsion of the nuclei, and a proton, or other positive

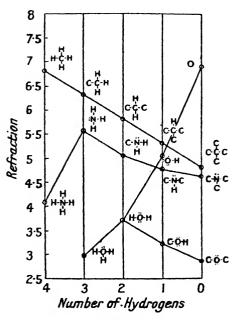


FIGURE IV.

(Graph from Fajans and Knorr, Ber., 1926, 59, 256; ct. C. P. Smyth, Dielectric Constant and Molecular Structure, p. 155. By permission of the Rheinhold Publishing Corporation, New York.)

group, from among those already attached, may be expelled, provided that an alternative acceptor molecule is at hand. Hydrogen compounds of low polarizability are consequently acidic in nature.

This structural relationship can be seen clearly from the above graph of ionic refractivities for the compounds of carbon, oxygen and nitrogen.

Induction of Acidity in Complex Hydrides.

It is a remarkable fact that, in complex hydrides of the type R—X—H, the readiness with which a proton is eliminated depends not only on the nature of the atom X, to which it is directly attached, but also on the radical R, although this radical does not play any direct part in the formation of the acid or base. Thus the very slight acidity of the hydroxyl radical in water is decreased still further when this radical is attached to an alkyl group as in the alcohols, but is increased substantially when it is attached to an aryl group as in the phenols. The acidity of the molecule therefore increases progressively in the series:

$$C_2H_5-O-H < H-O-H < C_6H_5-O-H$$
.

Even more remarkable is the fact that acidity can be induced in the inert molecule of methane by introducing suitable substituents. Thus Victor Meyer in 1872 ¹ observed (i) that nitro-paraffins, such as CH₃—NO₂, were capable of forming salts with the alkali metals ²; (ii) that similar acid-inducing properties were possessed by certain organic radicals, e.g. —CN, —CO·CH₃, —C₆H₅, in presence of which the adjacent atoms of hydrogen could often be displaced by metallic sodium, even when the hydrides were too feebly acidic to form salts with alkalis; and (iii) that the potential "acidity" of hydrogen atoms was also manifested, by reactions with halogens, nitrous acid, aldehydes, and diazonium salts, which would now be classified as kationoid reagents.

Thus the potential acidity induced in the methylene radical of ethyl acetoacetate by the contiguous —CO·CH₃ and —CO·OEt groups is shown most clearly by the formation

² These would now be formulated as containing the ion



¹ Ber., 1872, 5, 404.

of a sodium-derivative; but it is also indicated by the following reactions:

These acidifying groups can reinforce one another in a remarkable way. Thus, as Vorländer pointed out, the acidity of the carboxylic acids cannot be attributed either to the —OH group or to the —CO group alone, but is due to the combination of the two groups in the carboxyl radical, to which a special structure may perhaps be assigned, as in the formulæ

A similar reinforcement is also seen in the properties of ethyl acetoacetate, as recorded in the preceding paragraph, since the mobility of the hydrogen is almost inappreciable both in acetone and in ethyl acetate, which contain only one of the two acidifying groups.

Acylous and Basylous Radicals.

If has been found that the influence of acidifying groups is not confined to the vicinal atoms, but can extend through a chain of several atoms in molecules of sufficient size. Thus all organic nitro-acids have higher dissociation constants than their unsubstituted analogues. These influences must therefore be regarded as inherent characteristics of

¹ Ber., 1901, 34, 1632, 1637; Annalen, 1901, 320, 66, 69.

² Hantzsch, Ber., 1917, 50, 1438; Lowry, J. Chem. Soc., 1923, 123, 827. (Compare pages 36 and 395).

the radicals concerned, and can be used to classify them. Thus, Lapworth ¹ has suggested that "a group which tends inductively to lower the acidity of the molecule of which it forms a part" should be described as basylous, whilst "a group which has the opposite effect", e.g. a group which increases the acidity of a molecule, should be described as acylous. For example, the nitro group, —NO₂, may be called an acylous group, whilst the carboxylate anion, —COO⁻, may be called a basylous group. The acidifying radicals were defined by Victor Meyer in 1887 ² as "negative" groups, but this original terminology can easily lead to confusion when applied in conjunction with modern electronic conceptions and therefore will not be used again in this volume.

In general, acylous groups are radicals which are capable of forming free anions, e.g. Cl⁻, CN⁻, CH(COOEt)₂. Henrich,³ in 1899, pointed out that organic radicals of this type were all unsaturated, and, since "unsaturation" in an organic compound invariably produces an "exaltation" of its molecular refraction, one can say that acidity is induced by radicals of high polarizability. This conclusion might, at first sight, appear to be contradictory to that arrived at in the previous section, where it was shown that polarizable atoms were basic in character, i.e. tend themselves to combine with proton; but it is not so in reality, since it has been demonstrated that acylous groups usually promote a prototropic change (p. 304), as a result of which the mobile proton is attracted to the inducing group, and then unites with it, e.g.

$$Ph \cdot CH_1 \cdot \overset{+}{N} \overset{O}{\underset{O}{\longleftarrow}} \rightarrow Ph \overset{-}{\underset{C}{\longleftarrow}} H - \overset{+}{\underset{O}{\longleftarrow}} \overset{O}{\underset{O}{\longleftarrow}} Ph - CH = \overset{+}{\underset{N}{\longleftarrow}} \overset{\bar{O}}{\underset{O}{\longleftarrow}} H$$

¹ Mem. Manchester Phil. Soc., 1920, 64, No. 3, p. 2.

² Ber., 1887, 20, 534, 2994; 1888, 21, 1295, 1306, 1331, 1344, etc.

⁸ Ber., 1899, 32, 668, cf. "Theories of Organic Chemistry" chap. VIII.

Classification of Acylous and Basylous Radicals.

(a) Halogens. The mutual action between atoms of chlorine and hydrogen, which makes aqueous hydrogen chloride so strong an acid, persists in a lesser degree even when they are separated by one or more intervening atoms. All chloro-acids are therefore appreciably stronger than their unsubstituted analogues, as may be seen from the following tables:

TABLE I

Dissociation-constants	$(\times 10^{3})$	of	the Chloro-acetic	Acids at 25°
Acetic acid			CH, COOH	1⋅8
Chloroacetic acid.	•		CH,Cl·COOH	155
Dichloroacetic acid			CHČl ₁ ·COOH	5,000
Trichlocoacetic acid	1 .		CCI- COOH	30.000

TABLE II

Dissociation-constants (× 10⁸) of Monosubstituted Halogenated Acids at 25°

Acid.	CI	Br	Î	CI	Br	ī	CI	y Br	l I	Br	I	No Halogen.
Acetic . Propionic Butyric . Valeric .	155 147 139		75 	8·6 8·9	9.8	9·0 —	3.0	- 2·6 -	 2·3 	I·9	_ _ _ 1·7	1·8 1·4 1·5 1·6

- (i) From Table I it is obvious that the acylous effect of the halogens is cumulative, although perhaps less so than the above figures would suggest, since the coefficients of ionization are approximately proportional to the *square* roots of the dissociation-constants tabulated above.
- (ii) In the homologues of acetic acid the influence of the halogen diminishes progressively as its distance from the carboxylic hydrogen increases. In the ô-position, where it is separated from the hydrogen by five carbon atoms and one oxygen atom, the effect of the halogen is almost negligible. Langmuir 1 has suggested on theoretical grounds

¹ Chemical Reviews, 1929, **6**, 467.

that the acylous effect should decrease exponentially with increase in the distance separating the hydrogen and halogen atoms; this is in close agreement with the experimental data, the factor being approximately 2.7:I.

- (iii) In general it appears that chlorine has a slightly more powerful influence than bromine, which again has a greater inducing power than iodine, as would be expected on theoretical grounds.
- (b) Nitrogen Compounds. (i) The nitro-group is the most powerful of the acylous radicals. Thus, the presence of a nitro-group in an aliphatic compound is sufficient to make it soluble in caustic alkali, provided that the adjacent atom of carbon carries an atom of hydrogen, in the absence of which acidity would be no longer possible. Similarly, a nitro-group in an aromatic ring is sufficient to convert a phenol into a substance capable of decomposing sodium carbonate; conversely, it will make a nitrogenous base so weak that it will dissolve only in concentrated mineral acid. The quantitative effect of the nitro-group can be gauged from the following table:

TABLE III

Dissociation Constants of Phenols and Amines at 25°

Pho	enols.	Amines.				
o-Nitrophenol . m-Nitrophenol .	6.8 × · o - 8 3.9 × 10 - 8 7.0 × 10 - 8	p-Nitraniline .	1.0×10^{-14} 4.0×10^{-18}			

The powerful influence of the nitro-group may be attributed in the first place to the large nuclear charges of oxygen and nitrogen, which must both be regarded as acylous elements. In the nitro-group, however, this effect is reinforced by the presence of a semi-polar bond; this provides a permanent positive charge on the nitrogen, which is nearer to the acid group than is the negative charge on the oxygen; there is, therefore, a strong dipole tending to repel a proton, when

it seeks to recombine with the anion of the acid, as in the scheme:

$$0 \longrightarrow N \cdot C_{\mathbf{e}} H_{\mathbf{e}} \cdot OH \Rightarrow 0 \longrightarrow N \cdot C_{\mathbf{e}} H_{\mathbf{e}} \cdot O + H$$

(ii) The amino-group is basic, since it tends to accept a

proton. Nevertheless its action as a substituent on the strength of an acid is definitely acylous, as might be expected in view of the fact that the nuclear charge of nitrogen is greater than that of carbon. This acylous action is often obscured, however, by the formation of internal salts. Thus, in comparing the acidity of glycine, NH₂·CH₂·CO·OH, with that of acetic acid, it is necessary to bear in mind that the protons expelled from the carboxyl-group may be reattached to the amino-group, giving rise to an internal salt as indicated by the formula NH₃·CH₂·CO·O. It is, therefore, much more difficult to investigate the influence of an amino-group on the acidity of a carboxyl radical than to study the increase of acidity produced by a halogen. Direct evidence of the acylous character of the aminogroup was, however, provided by Vorländer in 1902 in an important paper "On the Nature of Radicals".1 Thus, from the data cited in Table IV, it can be seen that anilinoacetic acid is stronger than benzylacetic acid, although (as might be expected) it is not so strong as phenoxyacetic

TABLE IV

Dissociation Constants (× 10°) of Anilino-acids, etc.

C _e H _s ·O·CH _s ·COOH .		76	CH ₂ ·COOH	270
$C_6H_5\cdot NH\cdot CH_2\cdot COOH$	•	3.9		
C ₆ H ₅ ·CH ₂ ·CH ₃ ·COOH		2.3	$C_4H_4\cdot CH$ $CH_4\cdot COOH$ $CH_4\cdot COOH$	77
CH ₃ ·CH ₃ ·CH ₃ ·COOH	•	1.4	сн соон	
CH ₃ ·CH ₃ ·COOH	•	1.4	CH, COOH	4.7

¹ Annalen, 1902, 320, 99.

acid. Similarly, in the diacetic acids, the replacement of CH by N increases the strength of the acid.

Another example is provided by aspartic or aminosuccinic acid, which (unlike amino-acetic acid) is stronger than the acid from which it is derived, apparently because the neutralization of one carboxyl group by an α -aminogroup is so complete that its neutralizing action on the second carboxyl group is eliminated.

- (c) Aromatic Radicals. Aromatic radicals are generally classed as acylous groups in view of the fact that phenol is more acidic than methyl alcohol, and aniline is less basic than methylamine. This classification is supported by the fact that aromatic rings also increase the mobility of adjacent hydrogen atoms, e.g. phenylacetic acid is more reactive than acetic acid in the Perkin condensation, and desoxybenzoin, C₆H₅—CO·CH₂—C₆H₅, is more reactive than acetone. Similar properties are shown by other types of aromatic ring compounds, such as indene, fluorene and anthracene, which also contain reactive hydrogen atoms. This reactivity, however, is partly attributable to the presence in all aromatic rings of a mobile system of electrons, which can tolerate either a positive charge, as in the kation of triphenylmethyl chloride, CPh₃Cl, or a negative charge, as in the anion of the potassium derivative of pyrrol, C₄H₄NK.¹ In general, the effect of aromatic ring systems is best compared with that of unsaturated linkages, since no sharp distinction can be drawn between unsaturated systems, conjugated systems, semi-aromatic rings like
- (d) Alkyl and Alkylene Radicals. The dissociation constants of a series of saturated and unsaturated aliphatic acids are set out in Table V on page 211.

pentadiene, and true aromatic rings.

(i) Saturated Acids. Except on passing from formic to acetic acid, there is no dramatic change in the strength of these acids. The interpretation of the data, therefore, depends on choosing a conventional standard of reference

TABLE V

Dissociation Constants of Fatty Acids \times 10.

U	Insatura	ted Aci	ds.		
		αβ	βγ	γδ	δε
H-CH:CH-COO	H	5·60	• •	•	
CH ₃ ·CH:CH·CO	OH	2.0			
iso ,,		3.6			
C ₂ H ₅ ·CH:CH·CO	HOC	1.48	3.35	2.09	
α-Me ,,		1 · I	3.0	2.16	
β-Me ,,		0.73	2.38		
α-Et ,,		2.05	3.39		
C ₃ H ₇ ·CH:CH·CO	HOC	1.89	2.64	1.74	1.91
	Saturate	d Acid	s.		
H-COOH	21.4	C.H	.COOH	1	·43
CH₃·COOH	1·8	C,H	, COOH	I	.5
C₀H¸∙COOH	1.43		COOH		·6
C _a H ₇ ·COOH	1.5		11-COOF		·45

in relation to which the acylous or basylous character of the radical shall be judged. Thus if the hydrogen in formic acid were regarded as neutral, the methyl-radical in acetic acid must be strongly basylous, since acetic acid is much weaker than formic acid; but the strength of the higher homologues is almost constant and actually passes through a minimum at propionic acid. The growing hydrocarbon chain has therefore no important influence on the strength of the acid, and would provide a convenient standard of neutrality, relatively to which all other substituents, e.g. -Cl. -NO₂, -NH₂ (and even the replacement of the growing chain by hydrogen in formic acid) would be definitely acylous. In practice, however, it has become customary to take the methyl-group of acetic acid as a standard of neutrality. The replacement of hydrogen by methyl then weakens the acidity of the compound, and the alkyl group becomes the sole example of an uncharged 1 basylous radical, with the possible exception of radicals derived from atoms of lower atomic number, such as boron.

(ii) Unsaturated Acids. The dissociation constants of

¹ The carboxylate anion, —CO·O, is strongly basylous, for all dibasic acids have very small second dissociation constants (cf. p. 213).

unsaturated acids are of special interest on account of the fact that they pass through a maximum as the double bond is displaced along the hydrocarbon chain. Moreover, the hexenoic acids show indications of a definite alternation of strength in consecutive members of the series. It is also noteworthy that the homologues of acrylic acid exhibit a minimum when $R = C_2H_5$ in the acid R·CH:CH·COOH, just as the homologues of formic acid exhibit a minimum when $R = C_2H_5$ in the acid R·COOH. It is, therefore, difficult to resist the conclusion that the ethylenic group transmits to the carboxyl the influence of the growing hydrocarbon chain, but without itself counting as a part of that chain.¹

In general, the dissociation constants of unsaturated acids are appreciably higher than those of their saturated analogues. This is in accordance with the generalization of Henrich, already mentioned (p. 206), that unsaturation is conducive to the development of acidity; but one cannot stress too much the experimental fact that the acylous or basylous influence of any hydrocarbon radical is quite insignificant.

(e) Directly-attached Hydrogen Atoms. The selection of acetic acid as a standard of reference makes it necessary to regard formic acid as a compound in which methyl has been replaced by directly-attached hydrogen, which must therefore be regarded as a strongly acylous radical. The strength of formic acid, $K = 21.4 \times 10^{-5}$, as compared with acetic acid, $K = 1.8 \times 10^{-5}$, or carbonic acid, $K = 0.03 \times 10^{-5}$, is perhaps surprising. It finds a parallel, however, in the series H_3PO_4 , H_3PO_3 , H_3PO_2 , where there is a progressive increase in the equivalent conductivity of the acid as the content of oxygen is diminished. This might, at first sight, appear to be in contradiction to the conception of oxygen as an acidifying relement, but both series of observations can be explained

¹ Lowry, Phil. Mag., 1923, [vi] 46, 1013; Trans. Faraday Soc., 1923, 19, 500.

² Latimer and Rodebush, J.A.C.S., 1920, **42**, 1429.

in a simple way by means of the electronic theory, as follows:

It has already been indicated (p. 201) that hydrogen has a basylous action in the series HF, H₂O, H₃N. This depends on the progressive removal of protons from the central nucleus to the exterior of the octet, thus diminishing the restraining forces on the "lone pairs" of electrons which are still free, and endowing them with a higher polarizability. The data quoted above show that a still further reduction of acidity occurs when this directly-attached hydrogen atom is replaced by methyl or any other group containing at least one completed electronic shell.

On account of its minuteness the proton is able to penetrate into the electronic shell of another atom (p. 19) and, in part, act as an increment to the central nuclear charge in restraining the "lone pairs of electrons", whilst a larger group can only exert a force from the exterior.

Thus the influence of directly-attached hydrogen cannot be regarded as involving merely *induction* of reactivity, but depends on a more direct action upon the acidic group.

For that matter, there is always an intimate coupling of the valency forces of any two groups attached to the same atom, as in X—C—Y, which is not exerted between radicals separated by a chain of atoms.

Relationships Between Acidity and Molecular Configuration.

(a) Influence of Ionization on Acidity. It is well known that the univalent ion of a dibasic acid is much less easily ionized than the acid itself, since the second dissociation constant is always much smaller than the first.

Acid. Dissociation Products. Dissociation Constants.
$$H_3SO_4 \rightleftharpoons \overset{+}{H} + \overset{-}{HSO_4} \rightleftharpoons \overset{+}{2H} + \overset{-}{SO_4}$$
 0.45 0.017 $H_2CO_3 \rightleftharpoons \overset{+}{H} + \overset{-}{HCO_3} \rightleftharpoons \overset{+}{2H} + \overset{-}{CO_3}$ 3 × 10⁻⁷ 1.3 × 10⁻¹¹ Wegscheider 1 pointed out in 1895 that on statistical 1 Monatshefte, 1895, 16, 153.

grounds alone the second dissociation constant of a dibasic acid should be a quarter of that of the first dissociation constant, since (i) the probability of ionization of hydrogen from a dibasic acid molecule should be twice that of a similar molecule having only one ionizable hydrogen, and (ii) the probability of recombination of a hydrogen ion with a divalent anion should be twice as great as with a univalent anion. The difference between the dissociation constants of polybasic acids is, however, usually much greater than this statistical ratio. This discrepancy must be ascribed to the fact that, although unionized carboxyl groups, —COOH, or their close analogues such as —COOEt, —CONH₂, —SO₃H are all distinctly acylous, the anion formed in the

first stage of the dissociation, e.g. —CO·O, carries a local negative charge, which attracts the dissociated protons and is therefore to be regarded as a basylous group, since it is obviously more difficult to split off a proton from a negatively-charged ion than from a neutral molecule.¹

This attraction decreases, in accordance with the law of inverse squares, as the distance between the two acidic groups increases and has been made the basis of a quantitative method for calculating this distance, since N. Bjerrum in 1923 deduced values for the intramolecular distance between the protons of two carboxyl-groups from the deviations from the 4:1 rule of the ratio of first and second dissociation constants of dicarboxylic acids. He regarded the proton as unique, since it has a charge of but negligible dimensions, and can move under the action of electrical forces from carboxyl groups to any basic acceptor molecules. Hence its environment is conditioned by the energy levels in the vicinity of larger, less mobile atoms.

In a half-dissociated dibasic acid the direct effect of the field of the carboxylate anion, COO⁻, is much greater than any other electrical influence. The change ΔE in energy

¹ Ostwald, Zeit. phys. Chem., 1892, 9, 558; cf. Zeit. phys. Chem., 1889, 3, 415.

² Zeit. phys. Chem., 1923, 106, 219.

of environment of the proton which is produced by the presence of the other anion is given by the equation

$$\Delta E = \frac{Ne^2}{Dr},$$

where r is the distance between the carboxylate anion and the dissociable proton, and D is the dielectric constant of the medium. This change of energy is proportional to the logarithm of the resulting change of dissociation constant. The mathematical symbols

$$\log \frac{K_1}{K_2} - \log 4 = \frac{Ne^2}{RTDr}$$

and, substituting numerical values for the molecular constants, one gets

$$\log K_1 - \log K_2 - \log 4 = 3.1 \times 10^{-8}/r.$$

The way in which the dissociation constants of dibasic organic acids vary with increasing distance between the carboxyl groups is shown in the following table.

TABLE VI

Values calculated from Dissociation Constants at 25°

Acid.			$Log K_1$.	Log Ka.	$\log K_1 - \log K_1$	$n = \operatorname{Log} K_1 - \log_4.$
Oxalic			1.42	4.35	2.93	2.53
Malonic			2.80	5.69	2.98	2.29
Succinic			4.20	5.62	I · 42	0.82
Glutaric			4.32	5.50	1.18	o·58
Pimelic			4.49	(5.36)	(0.87)	(0.27)
Suberic			4.52	5.55	1.03	0.43
Azelaic	٠		4.60	5.56	0.96	0.36
Sebasic			4.62	5.60	I · O2	0.42

Using more accurate, recent data, derived from electrometric measurements, Gane and Ingold ¹ have further developed this estimation of intramolecular distance, and have shown that for the acids of the series HOOC— $(CH_2)_n$ —COOH the distance r separating the carboxyl groups is approximately r = (4.4 + 1.73n) Å.

The deviations from this mean equation for values of r are as follows:

¹ J. Chem. Soc., 1928, 1594.

There is consequently no indication that saturated hydrocarbon chains tend to coil round, as might be suggested from the ease with which homocyclic rings of five and six members are formed. Further, the calculated increment of distance per CH_1 group, $\delta r = 1.73$ Å, is greater than that deduced from crystal measurement, for which the calculated atomic distances are 1.54 Å for the carbon—carbon distance in diamond and 2.53 Å for the distance between corresponding alternate carbon atoms in the zigzag citains of the dibasic carboxylic acids.¹

Gane and Ingold suggest that the discrepancy between the two methods of calculation of the intramolecular distances must be attributed mainly to the effect of "electrostriction" (p. 250) in solution, which, by diminishing the effective pole strength of the — COO^- group would increase the apparent value of r.

A similar investigation of $\beta\beta$ -'dialkylglutaric acids affords support for the well-known theory of valency deflection of Thorpe and Ingold,² according to which any large substituent group attached to a carbon atom should decrease the angle between other valency directions, and vice versa.

TABLE VII

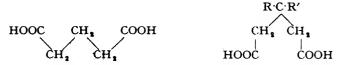
Substance.			7.*
Glutaric acid			9.22
β -Methylglutaric acid			2.27
β -n-Propylglutaric acid .			2.12
$\beta\beta'$ -Dimethyglutaric acid .			1.57
$\beta\beta'$ -Diethylglutaric acid .			1.02
ββ'-Di-n-Propylglutaric acid			1.01
CycloPentane-1: 1-diacetic acid			i ·40
CycloHexane-1: 1-diacetic acid			1.04
CycloHeptane-1: 1-diacetic acid			1.10
CycloHexane-1: 1-diacetic acid			1.04

¹ Caspari, J. Chem. Soc., 1926, 3235.

² J. Chem. Soc., 1915, 107, 1080; 1921, 119, 305; 1922, 121, 2676; 1922, 121, 1430; 1923, 123, 1206.

^{*} Gane and Ingold, J. Chem. Soc., 1928, 2267.

From these data it would appear as if glutaric acid itself possessed a zigzag carbon chain, but that the substituted acids had a coiled configuration:



Apart from this, the measurements of dissociation constants are in close agreement with the earlier deductions of Thorpe and Ingold, from observations of chemical reactivity.

In the case of the substituted malonic acids the apparent distances r between the two carboxyl groups, calculated by Bjerrum's method, is as follows:

$$R,R':H,H$$
 Me,H Et,H Pr_{α},H Pr_{β},H Me,Me Et,Et Pr_{α},Pr_{α} ! $r:1.54$ 1.64 1.50 1.47 1.37 1.45 0.71 0.64

These numbers are *not* in agreement with deductions from the hypothesis of valency deflection, which would place the acids in the following order:

r large H,H > Me,H > Et,H > Pr
$$_{\alpha}$$
,H >

$$Pr_{\beta}$$
, $H > Me$, $Me > Et$, $Et > Pr$, $Pr - r$ small

This discrepancy is attributed to the basylous character of the alkyl-radicals.

In the malonic acid series, the substituent alkyl groups are sufficiently near to the carboxyl groups for their basylous character to decrease the dissociation constants of the substituted acids, and thus to produce an apparent increase in the distance r.

Since this effect can scarcely be detected after transmission through more than one carbon atom in a saturated chain (cf. p. 211), it will be approximately equal for all the monosubstituted malonic acids and again about equal for all members of the group of disubstituted malonic acids.

The derived values of the distance r for the series of acids

r large Me, H > H, H > Et, H > Pr
$$_{\alpha}$$
, H > Me, Me > Pr $_{\beta}$, H > Et, Et > Pr, Pr r small

¹ Gane and Ingold, J. Chem. Soc., 1929, 1691.

therefore descend in the order that would be anticipated theoretically.

That the basylous influence of the alkyl groups affects principally the first dissociation constant K_1 , the second K_2 being controlled principally by the electrostatic effect of the carboxylate ion, —COO⁻, can be seen from the results of the dissociation constant measurements themselves.

TABLE VIII

Dissociation Constants of Substituted Malonic Acids

Groups.		$K_1 \times 10^4$	$K_3 \times 10^7$.	Groups.		$K_1 \times 10^4$.	$K_3 \times 10^7$
$\mathbf{A} \mathbf{B}$				A B			
H, H .		17.7	43.7	Me,H .		10.7	34.3
Et, H.		12.6	28·1	Pr_{α}, H .			20.8
Pr_{β} , H .		11.7	15.9	Me,Me .		8.27	15.3
Et, Et		62.3	0.590	\Pr_{α}		86.7	0 342

(b) Stereochemical Relationships. In considering the effect of the halogens upon acid strength it was pointed out that the acylous influence diminished progressively, in approximately exponential manner, with increase in distance between the halogen and the ionizable hydrogen (p. 207). This relation applies also to stereoisomers, and has frequently been used in determining their configurations. Thus, in the case of dibasic acids, it is found invariably that the cis-acid has a greater first dissociation constant, but a smaller second dissociation constant, than its transisomer, e.g.

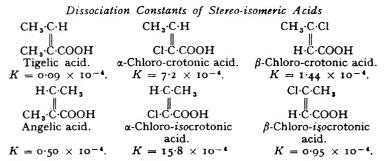
HC·CO·OH	HO-CO-CH
HC·CO·OH	HC·CO·OH
Maleic acid.	Fumaric acid.
$K_1 = 150 \times 10^{-4}$	$K_1 = 10 \times 10^{-4}$
$K_2 = 0.05 \times 10^{-5}$	$K_2 = 3.4 \times 10^{-8}$

Even among weakly acylous or basylous groups steric proximity is significant. For example, Hey 1 has called

¹ J. Chem. Soc., 1928, 2321.

attention to the necessity of considering the geometrical proximity of groups in the unsaturated monobasic acids which are set out below:

TABLE IX



The proximity of a CH₃ or halogen group to the carboxyl group in every case increases the acidic strength, though the effect is much less than for the mutual effect of two carboxyl groups.

Similarly, in the aromatic series, acids with an acylous substituent in the *ortho*-position are considerably stronger than their *meta* or *para* isomers, in which the effect of the substituent is of the same order as that of a γ - or δ -substituent in the aliphatic series.

Table X
Dissociation Constants of Aromatic Acids (× 104)

	200000000000000000000000000000000000000	00,0	3,000,000	٧,	11101110000	110000 (>	10,
	Acid.				Ortho.	Meta.	Para.
	Chlorobenzoic				13.0	1.5	0.93
	Bromobenzoic				14.5	1.37	0.59
	Hydroxybenzoic				10.0	0.83	0.29
_	Phthalic .				12.0	2.9	
•	Toluic .	•	•		1.2	0.56	0.43

(c) Unsaturation and Conjugation. In 1902 Flürscheim 1 called attention to the fact that "the $\alpha\beta$ and $\gamma\delta$ unsaturated acids are weaker than $\beta\gamma$ unsaturated acids, notwithstanding that in the $\alpha\beta$ -acid the negative double linking is nearest the carboxyl group". He regarded this as evidence that the unsaturated group exerts an opposite

¹ J. Prakt. Chem., 1902, 66, 321; 1909, 95, 718.

influence on alternate members of a saturated chain of atoms attached to it. A study of Table XI shows, however, that the replacement of a single bond by a double bond in the $\alpha\beta$ -position increases the strength of both propionic and butyric acids, decreases the strength of valeric acid, and again increases the strength of hexoic acid. Moreover, the variations of strength are comparable in magnitude with those produced by stereo-isomerism, e.g. in crotonic and isocrotonic acids, as shown in Table XI.

TABLE XI. Dissociation Constants (\times 10 $^{\circ}$) of Unsaturated Chloro-acids.

	Unsaturated Acids.		(iso)	Saturated Acids.	
β-Chloro-acid	CHCH=CH-COOH CHCCI=CH-COOH CHCH=CCI-COOH	14.4	9.5	CH, CH, CH, COOH CH, CHCI CH, COOH CH, CH, CHCI COOH	139 8·9 1·5

It seems, therefore, that the steric influences discussed in the previous section play a great part in the relatively small fluctuations that are produced by moving the double bond along the chain. Thus the maximum in the $\beta\gamma$ -position may perhaps be related to the maximum of stability in the γ -lactones; and the rather erratic alternations in the dissociation constants of the higher fatty acids may be associated with the zigzag structure of the hydrocarbon chain.

Conjugated systems, e.g. in aromatic hydrocarbons, frequently act as if they were amphoteric, and do not exert any very powerful acylous (or basylous) action; but they play a very important part in transmitting the inductive influence of acylous radicals to an acid group, over a much greater range of action than in saturated aliphatic systems.

¹ Cf. Kermack and Robinson, J. Chem. Soc., 1922, 121, 433; Robinson, Solvay Report, 1931; Ingold, Chemical Society Annual Reports, 1928, p. 121; Burton and Ingold, J. Chem. Soc., 1928, 904. See also p. 461.

CHAPTER X

THE REACTIVITY OF HALOGEN COMPOUNDS

Kationoid and Anionoid Reactions of the Halogens.

In the preceding chapter a number of aspects of the reactivity of hydrogen have been discussed; in the following paragraphs the influences affecting the reactivity of halogen atoms in organic substances will be traced in a similar way, with reference to reactions such as substitution, hydrolysis, reduction and the Friedel-Crafts synthesis. These reactions differ considerably both as regards the experimental conditions under which they take place, and as regards the velocity with which they proceed at normal temperatures; nevertheless, a consistent theoretical interpretation, based upon the accepted electronic structure of halogen atoms, can be given of the various types of chemical change in which they take part.

In the free state the atoms of the halogens carry an outer shell of seven electrons; but they tend to acquire a complete octet, either by becoming negative ions, e.g. Cl⁻, or by sharing a pair of electrons and forming a covalent linkage. The ionic form appears to be the more stable, since the halogens and their compounds readily yield halogen ions, whereas the halogen ions are comparatively inert. Hence, as a general rule, a covalent halide, on activation, breaks down into a negative halide ion and a positive organic radical:

$$CH_3 \cdot I \longrightarrow CH_3 + + I^- : C_6H_5 \cdot CO \cdot Cl \longrightarrow C_6H_5 \cdot CO + + Cl^-.$$

Under certain conditions, however, the rupture of the bond can take place in the opposite way, giving rise to a positive halogen ion. This occurs, for instance, in the process of reduction or inverse substitution, e.g.

The expulsion of a positively-charged halogen atom in these instances of inverse substitution is inferred from the fact that the atom of hydrogen which enters the molecule in place of the halogen may be obtained most easily from an acid, from which it separates with a positive charge.

Positively-charged halogen ions are also produced very readily by the polar activation of halogen molecules (p. 167). They are much more reactive than the negative halogen ions, since they exhibit a strong tendency to complete their depleted octets either by addition to an atom with an available lone pair of electrons, to form a covalent bond, or, alternatively, by acquiring the two extra electrons from an outside source, e.g. from a reducing agent (cf. p. 233). They can, therefore, be recognized by their power of liberating iodine from solutions of potassium iodide, or nitrogen from solutions of hydrazine.

$$\ddot{I} + \ddot{I} = I_2$$
: $2\ddot{C}IOH + N_2H_4 = N_2 + 2H_2O + 2HCI$.

"Inverse substitution" should, however, be regarded as an abnormal mode of reaction of organic halides, since a covalent linkage normally tends to break in such a way that the halogen atom acquires the two electrons of the bond, unless the halogen atom is attached to a radical which has a very powerful tendency to form a negative ion, e.g. $HC(COOEt)_2$.

Inverse Substitution.

(a) Halogen Compounds. The conditions under which inverse substitution takes place in halogen-compounds can be illustrated very clearly by the behaviour of the bromo-

derivatives of methane, where the action of water or of an alkali leads to hydrolysis in the mono- and di-derivatives and to inverse substitution in the tri- and tetra-derivatives.

Iodine is even more easily expelled as a positive radical by inverse substitution, as in the conversion of iodoform into methylene iodide,

$$CHI_3 + H_2O \rightarrow CH_2I_2 + HIO$$
,

and of methylene iodide into methyl iodide. The ease with which iodine can be expelled is indeed so great that it can be eliminated in an elementary form from the tetrahalogen derivatives without the intervention of an alkali, e.g.

$$\begin{array}{rcl} 2CCl_{3}I &=& I_{2} + C_{2}Cl_{6} \\ 2CBr_{3}I &=& I_{2} + C_{2}Br_{6} \\ 2CI_{4} &=& 2I_{2} + C_{2}I_{4}. \end{array}$$

It will be noticed that in the last action a second atom of iodine is split off from each carbon atom by a reversal of the usual procedure of addition of a halogen to an olefine.

(b) Nitro Compounds. The reduction of dibromocamphor to monobromocamphor by the action of alcoholic potassium hydroxide finds a parallel in the reduction of $\alpha\alpha'$ -bromonitrocamphor, which takes place under similar conditions, but much more violently:

$$2\text{NaOC}_2\text{H}_5 + \text{C}_8\text{H}_{14} \\ \downarrow \\ \text{CO} \\ + \text{NaBr} + \text{C}_2\text{H}_4\text{O} + \text{C}_2\text{H}_6\text{OH}.$$

In each case the "oxidizing power" of the "positive" halogen is probably expended in converting alcohol into aldehyde; but the greater vigour of the latter action may be accounted for by the fact that, in addition to bringing about a reduction of the bromo-compound by inverse substitution, the alkali also converts the reaction-product into a salt. The heat of neutralization of nitrocamphor by sodium ethoxide is therefore thrown into the process, and

causes the action to take place with almost explosive violence when carried out in concentrated solutions.

Burton and Kenner, in seeking to explain the elimination of a halogen from bromotrinitromethane, or of a nitrogroup from tetranitromethane

$$CBr(NO_2)_3 \longrightarrow CH(NO_2)_3$$

 $C(NO_2)_4 \longrightarrow CH(NO_2)_3$

concluded that "compounds which, by the replacement of an atom or radical by hydrogen, become tautomeric readily undergo inverse substitution under the influence of suitable reagents, among which sodium ethoxide and alcoholic potassium hydroxide are the most prominent". This explanation was contested by Macbeth,2 in view of the fact that the hydrogen of bromoform or of methylene iodide possesses no analogous mobility, and cannot in any sense be regarded as "tautomeric". The reduction of the bromomalonic esters and analogous compounds may therefore be more correctly attributed to the influence of the oxygen atoms of the α-carboxyl groups which "endow the halogen atoms with an electro-positive character".8 There is, in fact, an obvious analogy between the expulsion of a proton from an acid under the influence of acylous groups and of a "positive" halogen from a substituted hydrocarbon by a process of inverse substitution under the influence of similar groups. The "tendency to acquire a tautomeric hydrogen atom "which Burton and Kenner regard as "at least contributory to these reactions", is indeed only another way of describing a hydrogen atom which readily migrates in the form of a proton in the process of prototropic change (p. 300), and this migration is obviously a parallel phenomenon to the separation of a positivelycharged halogen from the same negatively-charged radical.

The preceding arguments may be summarized by stating that acylous groups, which are sufficiently powerful to

¹ J. Chem. Soc., 1922, 121, 489.

² Ibid., 892, 904, 1109, 1116 et seq.

⁸ Lapworth, Mem. Manchester Phil. Soc., 1920, p. 8; cf. Waters, J. Chem. Soc., 1933, 1556.

cause the easy liberation of protons from such stable linkages as $-CH_2$ — or $-NH_2$, can also cause the expulsion of positive halogen ions from the corresponding halides. This deduction may be illustrated by recalling the readiness with which inverse substitution takes place in the halogen derivatives of $\mathbf{1}:3$ diketones or ketonic acids, and of the sulphonamides. Lapworth (in a private note) has expressed this as a simple rule, thus:

"If X: (where: is an electron-pair) represents an *anion* with not too large free energy, then the halogen in X:hal will exhibit some 'positive' character and will readily be displaced by H; H:X is in fact an acid or proton donor."

This, of course, includes NH₃ and its derivatives (cf. Cl·NH₂) since NH₃ gives a fairly stable anion in Na·NH₂.

Reactivities of the Halogens.

The electrode discharge potentials of the four halogens follow the order of their atomic weights:

most stable anion F > Cl > Br > I least stable anion -1.359 - 1.066 - 0.536 volts.

The negative discharge potentials, therefore, decrease with increase of atomic weight, atomic dimensions or atomic polarizability (p. 257). This fact is capable of simple explanation, since the orbits of the valency electrons in fluorine and its compounds must be much closer to the central nucleus than in iodine and its compounds, where other completed electronic shells intervene. The valency electrons of the fluoride ion, therefore, move under the action of much stronger electrical fields, and consequently need much more energy of activation before they can be released to take part in chemical changes. Organic iodides should therefore be more reactive than similar derivatives of the other halogens, as is indeed found experimentally, e.g. Grignard reagents are prepared much more easily from alkyl or aryl iodides than from other halogen compounds.

The same data explain why the halogens displace each

¹ Lewis and Randall, "Thermodynamics", p. 433.

other from their ions in the order of their atomic weights. This displacement can be formulated in stages as follows:

(i)
$$Cl_{\frac{1}{2}} \rightarrow Cl^{+} + Cl^{-}$$

(ii) $Cl^{+} + I^{-} \rightarrow Cl^{-} + I^{+}$
(iii) $I^{+} + I^{-} \rightarrow I_{\frac{1}{2}}$

The essential stage of the process is shown in equation (ii). The loosely-held electrons of the iodide ion are there transferred to the depleted shell of the "positive chlorine" to make a more stable octet in the chloride ion. This process is a mere interchange of ionic charges and is obviously controlled by the difference between the potentials cited at the head of the section.

Induction of Reactivity in Organic Halides.

(a) Displacement of Chlorine by Iodine. In America, Conant and his colleagues ¹ have compared the reactivities of a large variety of organic chlorides by studying their interaction with potassium iodide in dry acetone.

$$RCI + KI = RI + KCI$$
.

This interchange of halogens was followed by titrating the residual iodide against potassium iodate in hydrochloric acid at o°, as represented by the equation

$$2KI + KIO_3 + 6HCl = 3KCl + 3ICl + 3H_2O$$
.

Free iodine is formed at first and the end-point was shown by the decolorization of a chloroform extract when the iodine was finally converted into iodine monochloride. This method of studying the reactivity of organic chlorides has the following advantages:

- (i) The solvent is inert to all types of organic chlorides, so that a wide range of substances can be examined.
- (ii) The reagent consists of simple iodide ions, since the reaction velocity was not affected appreciably when lithium or sodium iodide was used in place of potassium iodide.²

¹ J.A.C.S., 1924, **46**, 232; 1925, **47**, 476, 488; 1926, **48**, 2745; 1928, **50**, 2446; 1929, **51**, 3409.

² Conant and Hussey, J.A.C.S., 1925, 47, 482.

The reaction velocities in acetone at 50° of different series of compounds of the type $R \cdot (CH_2)_n \cdot Cl$ relatively to that of butyl chloride as unity, are set out in the following table.

TABLE I ¹
Reactivity of Halides $R(CH_1)_{n}$ ·Cl towards Iodide Ion

R	n = 0	I	2	3	4	5	6	7
н			2.52	10.8	1.00	1.35	1.30	1.25
Ph .	-	197	1.13	1.72			1.46	
	ca. 2,200	ca. 10 ⁸	86∙7	372				
EtO-OC	42	1,720	1.61	1.65	1.35			
Ph·S·	-		0.56	2.74				
Me⋅S・ .			1.52	2.52				

The theoretical significance of these measurements can only be gauged after the mechanism of the displacement process has been considered (see p. 232), but mere inspection shows that the induction and repression of reactivity in the halogens has similar characteristics to the induction and depression of acidity in the analogous compounds of hydrogen. Thus—

- (i) The activating or deactivating effect of the group R-decreases very rapidly as —CH₂— groups are inserted between it and the point of reaction within the molecule, exactly as in the case of the acids.
- (ii) In a wide range of alkyl chlorides, the differences in reaction velocity are small and irregular, like the variations in strengths of the homologues of acetic acid (p. 211); but secondary chlorides are much less reactive than primary chlorides.
- (iii) In general, the reactivity with potassium iodide of compounds of the type $R\cdot CH_2\cdot Cl$ diminishes in the following order:

$$Ph \cdot CO_{\bullet} > Me \cdot CO > CN > EtO \cdot OC > MeO > CH_{\bullet}CO > Ph > CH_{\bullet}:CH > C_{\bullet}H_{\bullet},$$

where it is obvious that the more acylous radicals precede those which are less acylous or more basylous.

- (b) Hydrolysis. In Holland, S. J. C. Olivier has carried out an extensive series of investigations of the hydrolysis of alkyl and acyl chlorides in aqueous alcohol and acetone,
- ¹ The table is taken from the Annual Reports of the Chemical Society for 1929.

whilst in this country a similar but less complete series of determinations was made by the late J. B. Shoesmith. Both investigators applied their results, with considerable success, to the prediction of the manner of substitution in the aromatic ring, since it appears that the reactivities of substituted aromatic halides can be correlated with ease of nuclear substitution. This aspect of their work, however, is only secondary in importance to the main facts that they have disclosed.

Both series of observations may be viewed as extensions of an important paper by Lapworth and Shoesmith 1 who showed, in 1922, that the ease of hydrolysis of the three isomeric methoxy-benzyl bromides exhibited an alternation, the ortho compound being more reactive than benzyl bromide, the meta compound less reactive, and the para compound much more reactive again. This work was, at the time, discussed widely, for it was considered to provide valuable support for the original theory of "Alternate Polarities", as formulated by Lapworth in 1920. However, further investigations have shown that the transmission of reactivity to substituted benzyl halides is of a much more complex character than was at first imagined, and has to be analysed into effects due to at least three factors. This can be shown by consideration of the following tables of results, compiled from the papers of the various authors.

Table II

Relative Velocities of Hydrolysis of Substituted Benzyl Bromides 2

Substituent.			Kortho Kunauba	Kmeta Kunsubs	Kpora Kunsubs,	Temperature.
CH ₃ O			49.6	0.70	50	60
CH_3 .			3.23	1.15	4.44	60
F.			0.42	0.30	1.36	60
\mathbf{F} .	•		0.46	0:33	1.84	76
Cl .			0.39	0.33	0.83	76
Br .			0.39	0.33	0.66	76
NO.			0.13	0.12	0.13	76
COOH	•	•		0.33	0.26	76

¹ J. Chem. Soc., 1922, **121**, 1391.

² Calculated from Shoesmith and Slater, J. Chem. Soc., 1924, 125, 2278, and 1926, 214; Shoesmith and Connor, J. Chem. Soc., 1927, 1769.

Table III

Relative Velocities of Hydrolysis of Substituted Benzyl Chlorides 1

Substituent.		Kortho Kunsubs. 4·84	Kmeta Kunsubs. 1.39	Kpara Kunauba. 10.6	Temperature.
		4.95	1.30	9.4	30
Cl .		0.355	0.237	0.62	83
			0.137	o·468	30
Br .	•,	0.286	0.215	0.50	83
		0.213	0.132	0.373	30
I.		0.287	0.201	0.477	83
		0.224	0.137	0.373	30
NO_2		0.084	0.090	0.074	83
		0.047	0.057	0.044	30
COOH			0.235	0.165	83
			0.170	0.100	30

For disubstituted compounds, Olivier found 2 that reaction velocities could, as a general rule, be computed from the data for the mono-substituted compounds by using a multiplication equation of the following type:

$$K_{3:4\text{-disubs.}}/K_{\text{unsubs.}} = (K_{3\text{-subs.}}/K_{\text{unsubs.}}) \ \times \ (K_{4\text{-subs.}}/K_{\text{unsubs.}}).$$

A similar equation to this had been applied successfully to the computation of the dissociation constants of disubstituted organic acids by Ostwald, as long ago as 1889,³ and was even then cited as evidence to suggest that substituent groups exerted their inducing influence upon reaction velocities independently of one another.

The two sets of observations quoted above are in accordance with one another, and are by no means the only recorded observations of reaction velocity measurements of this same type. They are remarkable in that in no simple instance is the value for the *meta* compound intermediate between those of the *ortho* and *para* compounds.

¹ Calculated from Olivier, Rec. trav. chim., 1922, **41**, 646; 1923. **42**, 516; cf. Chemical Society Annual Reports, 1927, p. 155.

² Rec. trav. chim., 1926, 45, 296, 455.

³ Zeit. phys. Chem., 1889, 3, 385.

⁴ Cf. Norris and Banta, J.A.C.S., 1928, 50, 1804.

230 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

On extending his researches from alkyl halides to acyl halides, however, Olivier found that diametrically opposite effects were produced, and his experimental results have been confirmed by other investigators. Thus, whilst the nitrobenzyl chlorides are from II to 20 times less reactive than benzyl chloride itself, the nitrobenzoyl chlorides are from 2 to II times more reactive than benzoyl chloride.

TABLE IV

Relative Velocities of Hydrolysis of Substituted Benzoyl Chlorides 1

			Kortho	Kmets	Kpara
Substituent.			Kunsubs	Kunsubs.	Kunsubs.
NO2.			1.77	6.82	11.54
Cl .			1.46	1.46	0.85
Br .			1.35	1.69	0.92
CH_3 .			large	1.46	2.88

TABLE V

Velocity of Esterification of Isopropyl Alcohol with Substituted Benzoyl
Chlorides 2

Acid Chloride.							Kaube./Kuna iba.
p-Chlorobenzoyl	•	•	•	•	•	•	1.53
p-Bromobenzoyl							1.78
p-Iodobenzoyl .							1.26
p-Methylbenzoyl							0.64
p-Nitrobenzoyl.	•						10.0
o-Nitrobenzoyl .							1.90

Consideration of the contrast between the hydrolysis of the alkyl and acyl halides will be deferred to the following section, dealing with the actual mechanisms of the reactions. One fact worth notice at this point, however, is that "steric" influences, which are not appreciable in the reactions of the alkyl halides, become of importance in those of the acyl halides.

(c) The Friedel-Crafts Reaction. There are many indications that the Friedel-Crafts synthesis of ketones

¹ Calculated from Berger and Olivier, Rec. trav. chim., 1927, 46, 517.

² Norris and Gregory, J.A.C.S., 1928, **50**, 1813.

³ Olivier, Rec. trav. chim., 1929, 48, 229.

⁴ Darzens, Compt. rend., 1910, 150, 707; Wieland and Bettag, Ber., 1922, 55, 2246.

from acyl chlorides and aromatic compounds depends on an initial addition of the acyl halide to a double bond followed by elimination of a halogen acid, thus:

This conclusion is supported by the fact that the reaction is not confined to aromatic substances, but can be applied also to olefines such as isobutylene 1 or cyclohexene.2 The rôle of aluminium chloride is interpreted by regarding the process as a catalysed reaction, in which aromatic molecules interact with molecular addition-compounds of acyl halides with aluminium chloride. These addition-compounds can be represented by the general formula [R·CO·Cl, AlCl₂] 3; but Meerwein 4 and W. Hückel 5 regard them as co-ordination-compounds, in which ionization of the organic chloride is indicated by transferring an atom of chlorine to a complex anion, thus:

$$(R\cdot CO)+(AlCl_4)-$$
 compare $H(B F_4)$

At least one molecular proportion of aluminium chloride

¹ Kondatow, Bull. Soc. Chim., 1892, (iii), 7, 576.

² Darzens, Compt. rend., 1910, 150, 707; Wieland and Bettag, Ber., 1922, 55, 2246.

³ Boeseken, Rec. trav. chim., 1900, 19, 19; 1901, 20, 102 et seq., to 1913, 32, 184; Olivier, Rec. trav. chim., 1914, 33, 91; 1915, 34, 109, 166; 1918, 37, 205.

⁴ Annalen, 1927, 455, 227; Ber., 1928, 61, 1840.

⁵ "Theoretische Grundlagen der Organischen Chemie", 1931, Vol. I, p. 351.

must be used per molecule of acyl halide, but any excess over this amount acts as a powerful catalyst also.

Olivier 1 has shown that under standardized conditions the order of reactivity of different acyl halides is the same as in hydrolysis of the substituted benzyl chlorides. Thus, the reactivity is increased by groups such as —CH₃ but is decreased by halogens or nitro-groups.

(d) Reduction of Alkyl Halides. In the initial investigation by Lapworth and Shoesmith ² of the reactivity of the three methoxy-benzyl bromides, it was shown that the replacement of bromine by hydroxyl (hydrolysis with alkali) and by hydrogen (reduction with hydriodic acid) were inverse processes, the one reaction being accelerated by influences which retarded the other and vice versa. Thus the halogen-substituted benzyl bromides are all more easily reduced than benzyl bromide itself, though they are hydrolysed less easily; and the rule appears to be a general one, as can be seen by comparing Table VI with Table II, (p. 228) both of which are compiled from one set of experimental results.

Table VI
Reduction of Substituted Benzyl Bromides with Hydriodic Acid

Substituent.	Kortho Kunnubs.	Kineta Kunsubs.	Kpara Kunsubs.	Temperature.
\mathbf{F} .	. 1.35	1.10	0.90	1010
Cl .	. 1.85	1.2	1.55	1010
\mathbf{Br} .	. 1.8	1.15	1.45	1010
COOH		0.60	1.66	110°

Mechanism of the Polar Reactions of Organic Halides.

Lapworth and Shoesmith found that acylous groups *inhibit* the liberation of negative halogen ions in the hydrolysis of alkyl halides, but *promote* the liberation of positive halogens in the process of reduction or "inverse substitution". This result might be expected, since when a positive halogen ion is formed by inverse substitution, the

¹ Rec trav. chim., 1922, **41**, 301; 1926, **45**, 710; 1927, **46**, 605. ² J. Chem. Soc., 1922, **121**, 1391.

halogen leaves both electrons of the bond behind, just as the proton does when an acid is ionized. On the other hand, when a negative halogen ion is formed from an alkyl halide, the two electrons of the bond are retained by the halogen and an opposite effect might be expected. Acylous groups, which assist in the separation of hydrogen ions in the form of protons, therefore promote the separation of positive halogens, but impede the separation of negative halide ions from alkyl chlorides.

This observation is of fundamental importance in the theory of polar reactions. Thus (a) It provides ample justification for the view that the inductive influence of substituents (which affect reaction-velocity without themselves taking part in the reaction) is in origin essentially electrical, as will be indicated in fuller detail in the following chapter. (b) The sign (i.e. acceleration or retardation) of the induction by substituent groups of reactivity in the organic halides becomes diagnostic in the determination of their reaction mechanisms. Thus Lapworth and Shoesmith inferred that in hydrolysis the halogen atoms separated as negative ions, since this process was inhibited by acylous groups (Cl. NO₂, etc.)

$$Ph \cdot CH_2 \cdot Br + OH^- \rightarrow Ph \cdot CH_2 \cdot OH + Br^-$$
.

On the other hand, reduction was attributed to the separation of positive ions of the halogens, since it was accelerated by acylous groups. The transference of electrical charges in reduction can be formulated as follows:

(i)
$$Ph \cdot CH_2 \cdot Br + H^+ \longrightarrow Ph \cdot CH_3 + Br^+$$

(ii) $Br^+ + 2\overline{1}^- \longrightarrow Br^- + I_2$.

This reaction is therefore governed by the decrease of total free energy (i) on converting a hydrogen ion into covalently-bound hydrogen, and (ii) on reversing the polarity of the active Br⁺ ion by means of iodide ions, which easily give up their negative charges and yield neutral molecules of iodine.

(c) It appears that the sign of ionization of a —C—Cl

234 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

bond depends on the sign of the ion by which it is attacked, e.g.

$$\vdots \overset{+}{C} + \overset{-}{Cl} \overset{-}{\longleftrightarrow} \vdots \overset{-}{C} + \overset{+}{Cl}.$$

This point of view has been expressed by C. K. and E. H. Ingold 1 in the following terms:

"The primary condition affecting the reactivity of the halogen substituents themselves is not so much the average, or actual, electrical condition of these atoms, but rather the extent to which they are capable of temporary electrical polarization by the external fields of the reagent molecules; in other words, the fundamental condition is the magnitude of the restoring forces brought into play when the electrons of the carbonhalogen linking are displaced in either direction. . . . seems a reasonable assumption that the relative displaceability of the shared electrons of the four halogens in corresponding halides will run parallel with that of the unshared halogen electrons, and there is an accumulation of evidence that whilst the unshared electrons which combine most readily with a positive charge are those of iodine (example: formation of iodonium salts), the electrons of the outer shell of iodine are also those which are most readily displaced by a combining

It is but a short step to the inference that the shared electrons in corresponding halogeno-hydrocarbons are most easily displaced in both directions (i.e. are most easily polarizable) in the case of the iodo derivative and least easily in the case of the fluoride. Thus the probable order of decreasing restoring forces is F > Cl > Br > I (cf. pp. 225-6) and this is the sequence required to account for the observations relating to the side-chain reactivity of the benzyl halides."

This point of view may be summarized by saying that the capability of reaction of any halide depends upon the external electrical fields of the reagent used as well as upon the internal electrical fields, due both to the actual electronic system of the carbon-halogen linkage and the induced electrical fields of any substituent groupings. It is then possible to provide a self-consistent explanation, not only of the hydrolysis and reduction of alkyl halides, but of all

other published work on the reaction-velocities of organic halides, ¹ on the assumption that the polar character of the reaction is attributed to the character of the reagent and not only to the organic compound itself. Some of these other reactions, for which velocity-figures have been determined are discussed in the following paragraphs.

(a) Hydrolysis of Alkyl and Acyl Chlorides. From a study of the hydrolysis of the substituted benzyl halides Olivier and Berger ² came to the conclusion that the determining ³ polar factor in the reaction was the formation of negative halide ions, followed by attachment of OH⁻ from the reaction mixture, thus:

(i)
$$Ph \cdot CH_2 \cdot Cl \rightarrow Ph \cdot CH_2^+ + Cl^-$$
 (slow)

(ii)
$$Ph \cdot CH_2^+ + OH^- \rightarrow Ph \cdot CH_2OH$$
 (rapid).

A similar mechanism may be postulated for the Friedel-Crafts reaction where it must be assumed that *ions* of the acyl halide add on to the double bond of the aromatic ring (pp. 230-232).

On the other hand, they consider that the primary reaction in the hydrolysis of acyl chlorides is the addition of water to the molecule of the acid chloride, followed by a rapid breakdown of the addition-product, thus:

(iii)
$$Ph \cdot CO \cdot Cl + H_{\bullet}O \rightarrow Ph - C \cdot OH \text{ (slow)}$$

(iv) $Ph - C \cdot OH \rightarrow Ph \cdot CO \cdot OH + HCl \text{ (rapid)}.$

Since, however, the hydrolysis of esters has been shown to depend on the addition of ions and not of neutral molecules (see Chapter XII), the above reaction can be represented more clearly as depending on an addition of a hydroxyl

¹ Cf. Bennett and Berry, J. Chem. Soc., 1927, 1676.

² Rec. trav. chim., 1926, 45, 712; 1927, 46, 516, 605; cf. van Duin, Rec. trav. chim., 1927, 46, 259; Chem. Weekblad, 1925, 22, 146.

³ It is one of many factors, cf. chap. XI, pp. 242-252.

ion to the carbonyl group (just as a cyanogen ion is added to it in the cyanhydrin reaction), followed by decomposition of the complex anion so formed ¹

$$(v) \ \, \text{Ph-C} \overset{\text{Cl}}{ \bigcirc_{\text{O}}} + \text{OP}^{-} \longrightarrow \text{Ph-C} \overset{\text{Cl}}{ \bigcirc_{\text{OH}}} \text{ (slow)}$$

(vi)
$$Ph$$
— C \xrightarrow{Cl} \longrightarrow Ph - C \xrightarrow{O} + Cl $^-$ (rapid).

The two reactions (i) and (v) are, however, opposite in polar character, since in reaction (i) the organic radical gives up a pair of electrons to the halogen atom, whilst in reaction (v) it accepts a pair of electrons from the hydroxyl ion. Any polar substituent which tends to expel electrons would therefore facilitate reaction (i) and inhibit (v); whilst groups which tend to accept electrons would inhibit (i) and facilitate (v) so that the two reactions would be affected in inverse directions. Thus acylous groups, which (by definition) increase the ease of ionization of an acid, and therefore favour the acceptance of an electron-pair by the negative ion which is left when the acid dissociates, should promote the hydrolysis of an acyl halide but not of an alkyl halide, as is actually the case.

(b) Interchange of Halogens. The hydrolysis of an alkyl halide with water or dilute alkali, and its reaction with potassium iodide are (like the hydrolysis of alkyl and acyl halides) reciprocal processes as regards their reaction mechanism, since the factors which assist the one retard the other, as can be seen from the data in the table opposite.

Since the displacement of chlorine by iodine in the alkyl chlorides is influenced by substituents in the same way as the displacement of chlorine by hydroxyl in the hydrolysis

¹ Berger and Olivier have shown however that the hydroxyl ion can be provided by a neutral water *molecule*, since the addition of dilute sulphuric acid does not decrease the rate of hydrolysis of acyl chlorides (see Berger and Olivier, *Rec. trav. chim.*, 1927, **46**, 618, 861). Compare the mechanism for ester hydrolysis that is suggested on pages 269–271.

TABLE VII

Reactivity of Halides 1

			Relative reactivity with	
Substance.			KI at 50°.	Water at 30°.
Ph·CH₂·Cl			1.00	1.00
$o\text{-NO}_2\cdot C_6H_4\cdot CH_2\cdot Cl$.			9.15	0.047
$m \cdot NO_2 \cdot C_6 H_4 \cdot CH_2 \cdot Cl$			3·96	0.057
$p - NO_2 \cdot C_6 H_4 \cdot CH_2 \cdot Cl$.			6.95	0.044
			KI at 55°.	Alkali at 75°.
$Ph \cdot S \cdot C_2H_4 \cdot Cl$			1.00	1.00
o-NO2·C6H4·S·C2H4·Cl			1.08	0.053
$m - NO_2 \cdot C_6 H_4 \cdot S \cdot C_2 H_4 \cdot Cl$			2.09	o∙o83
p-NO ₂ ·C ₆ H ₄ ·S·C ₂ H ₄ ·Cl			2.62	0.052
o-NO2·C6H4·CH2·S·C2H4	·Cl		2.36	
Ph·S·C ₃ H ₆ ·Cl			3.52	0.0031
p-NO2·C6H4·S·C3H6·Cl			9.19	0.0010

of acyl chlorides, Bennett and Berry ¹ have suggested that the process is again governed by a slow addition-reaction. They therefore postulate an initial addition of an iodide ion from the reagent to the organic chloride, followed by a rapid ejection of chloride ion from the complex so resulting:

(vii)
$$R \cdot CH_2 \cdot Cl + I^- \rightarrow R \cdot CH_2 \cdot \cdot \cdot \cdot I^-$$

(viii) $R \cdot CH_2 \cdot \cdot \cdot \cdot I^- \rightarrow R \cdot CH_3 - I + Cl^-$

The absence of the carbonyl group, however, makes this ad hoc explanation very unconvincing, since there is no obvious reason why an iodide ion (and not a hydroxyl ion) should be attracted to the completely saturated alkyl chloride, unless we fall back on a remote analogy with the formation of polyiodides or the co-ordination of a halogen ion with hydrogen in the acid fluorides. This reaction, however, receives further consideration in Chapter XI, in which it is shown that the resultant effect of several polar factors has to be discerned before any reaction mechanism can adequately be explained.²

¹ Bennett and Berry, J. Chem. Soc., 1927, 1676.

² See also Baddeley and Bennett, J. Chem. Soc., 1933, 261; 1935, 1819; Bennett and Jones, J. Chem. Soc., 1935, 1815, 1827; Olivier and Weber, Rec. trav. chim., 1934, 53, 869, 891.

CHAPTER XI

GENERAL POLARITY

Definition of General Polarity.

In the two preceding chapters it has been shown that the reactivity of hydrogen and of the halogens is influenced in a remarkable degree by radicals (e.g. halogens, alkyl and acyl radicals, nitro-, amino-, carbonyl- and carboxylgroups) which do not play any direct part in the reactions in question, since they are not affected by the reagents by which the reactions are brought about. This influence of one radical upon the reactivity of another radical in the same molecule may be either positive or negative in character. It was described by Flürscheim in 1902 as a manifestation of a "polar factor".

The type of polarity which has been described in the preceding chapters is characterized by the facts (a) that the influence of a given radical decreases progressively as it becomes more remote from the scene of action; and (b) that the effects which it produces remain constant in sign. For example, the replacement by chlorine of a hydrogen atom in the hydrocarbon chain of a fatty acid always brings about an increase of acidity (cf. p. 207). Lapworth has used the term general polarity to describe this specific type of influence, and, since 1920, his term has been given a more precise significance than was previously accorded to the use of the word "polarity" in organic chemistry. Thus, general polarity is regarded as an inherent property of a given radical; it must, therefore, operate

¹ Flürscheim, J. prakt. Chem., 1902, 66, 321; 1907, 76, 165, 185, 200; Ber., 1906, 39, 2016; J. Chem. Soc., 1910, 97, 85; 1926, 1562.

² Mem. Manchester Phil. Soc., 1921, 64, 13.

both in "normal" and in "activated" molecules, though the effects produced in the resting and in the activated states of the molecule may be different. Again, the general polarity of a given radical may manifest itself in many types of compounds, e.g. the general polarity of the nitrogroup increases the dissociation of nitro-acids, makes the methylene-groups of the nitro-paraffins reactive, promotes the "inverse substitution" of their halogen derivatives, decreases the ease of hydrolysis of alkyl halides and the basicity of amines, and so on throughout the whole range of reactions of nitro-compounds. The uniform sign of these effects, and their progressive decrease with increasing distance suffice to characterize them as manifestations of "general polarity".

Electrical Character of General Polarity. The ionization of an acid involves the development of a dipole between the ions, e.g.

$$CIH + OH_1 \rightleftharpoons CI + OH_3$$
.

It is therefore promoted by the varying ¹ field of an electric dipole acting in the same direction, and is inhibited by a dipole acting in the opposite direction. Substituents in which the dipole moment is directed away from the carboxyl group are therefore acylous, whereas those in which the dipole moment is directed towards the carbonyl group are basylous, e.g.

Acylous
$$Cl \cdot CH_1 \cdot CO \cdot O \cdot H + OH_2 \rightleftharpoons Cl \cdot CH_2 \cdot CO \cdot \overline{O} + OH_3$$

Basylous $CH_3 \cdot CH_2 \cdot CO \cdot O \cdot H + OH_2 \rightleftharpoons CH_3 \cdot CH_2 \cdot CO \cdot \overline{O} + OH_3$

Thus the nitro-group, which has a very strong dipole moment exerted away from the point of attachment

an alkyl radical is feebly basylous.

The electrical fields due to ions are still more powerful;

¹ Separation of electrical charges is not brought about by a uniform electrical field. A potential gradient is necessary.

thus an ammonium ion, in H_3N — $(CH_2)_n$ —COOH, is a powerful acylous radical, whilst a carboxylate anion, in

OOC—(CH₂)—COOH is the strongest of basylous radicals. The close parallelism between the effects of acylous groups on the separation of protons from acids and on the separation of positive halogens from alkyl halides then justifies the view that all the effects which are classified under the heading of general polarity may be ascribed to the influence of an electrical field of force set up by dipoles or ionized atoms.

Polar effects, however, can influence both the polar activation of covalent linkages in a molecule and also the ease of attack of a compound by the reactant molecules or ions around it. Thus, when an ionizable reagent is added to an olefine, a double bond in the molecule is activated to a semi-polar bond, which then reacts with the ions of the reagent. In this process the activation of the double bond and the ionization of the reagent may be affected by general polar influences operating between the two molecules. Consequently, it is found that the same olefine may yield different proportions of the various isomers when interacting with different additive reagents, or when it is in a different environment. This may be instanced by the following figures which have been quoted already (p. 179).

Addition to Propylene $CH_3 \cdot CH = CH_2$ HI gives $CH_3 \cdot CHI \cdot CH_3 = 99\frac{3}{4}\% = CH_3 \cdot CH_2 \cdot CH_2I = \frac{1}{2}\%$ ICI = 0.00, $CH_3 \cdot CHCI \cdot CH_2I = 75\% = CH_3 \cdot CHI \cdot CH_2CI = 25\%$ BrCI = 0.00, $CH_3 \cdot CHCI \cdot CH_2 \cdot Br = 58\% = CH_3 \cdot CHBr \cdot CH_2CI = 42\%$

The addition-reaction cannot, therefore, involve merely the addition of completely ionized reagents to a double bond which has been activated in a manner depending upon its own configuration and energy content alone, but must involve the action of other electrical polarization forces, traceable perhaps to dipole association (p. 64).

Thus, according to both Robinson 1 and Ingold,2 before an actual chemical change takes place, the polar fields of

¹ Solvay Report, 1931 (in which earlier papers are cited).

² J. Chem. Soc., 1931, 2354, 2746.

two interacting molecules must affect, and probably augment, each other by a process of "electrostriction" as the two molecules approach each other. This can be illustrated by means of a diagram (due to Ingold) which suggests how electron movements that can result in polar activation (indicated by the curved arrows) may tend mutually to promote each other by the approach in correct orientation of the reacting molecules.

As G. N. Lewis 2 originally conceived, one may consider that general polarity operates throughout the whole of a molecule, and upon vicinal molecules or ions, by a mechanism of electrostatic induction. Whenever a "polar group" can induce in a bond a dipole in the sense required for the chemical activation characteristic of it, it will increase the reactivity of this bond and vice versa. Thus p-nitrophenol is a stronger acid than phenol, whilst p-nitraniline is a weaker base than aniline, because the acylous nitro-group will induce an electrostatic displacement,3 or, in other terms, induce an "electron drift", towards the nitrogen atom, which will favour the release of proton from OH but will prevent the gain of proton by NH₂.

$$\vec{O} = \vec{N}$$
 $\vec{O} = \vec{N}$
 $\vec{O} = \vec{N}$

Electron displacement induced.

Electropositive field (diminishing with distance) attractive to anions repulsive to kations (e.g. proton).

¹ See p. 383 for an explanation of this symbolism.

^{2&}quot; Valence", 1923, pp. 84, 143.

³ Cf. Lucas, J.A.C.S., 1924, 46, 2475.

Lapworth, however, has urged that, although a given substituent, Y, usually has an "inherent" electron-attractive or repulsive power, it is not logical to suppose that its net effect on the electrons of the carbon atom to which it is directly attached in Y-CHR'R" is always in the same sense as its net effect on the electrons of more remote atoms. This effect is essentially a change in the electrical field around the molecule, and is equivalent to the superposition on the original field of a new external field due to the substituent. This new external field normally increases in intensity as the substituent is moved nearer to the scene of action in the molecule; but is a vector quantity, and, in Lapworth's opinion, the superimposed field may in some cases be so complex that neither its intensity nor its sign at one part of a molecule could be deduced from a knowledge of the distribution of the field at other points.

Sign and Symbolism of General Polarity.

Some confusion has arisen in reference to the nomenclature and symbolism used in describing the phenomenon of general polarity. Acylous groups are clearly electronattracting, since they include positively-charged radicals

such as $-NH_3$. Ingold ² therefore describes them as "electron sinks" and represents them by the symbol -I, in accordance with the general tradition, dating from Victor Meyer, which describes them as "negative" radicals.

Conversely, basylous groups, such as —CO·O, are clearly electron-repelling. They are therefore described by Ingold as "electron sources" and represented by the symbol + I, in accordance with the tradition which describes them as "positive". Since, however, the acylous groups include positively-charged atoms, Robinson 3 has suggested that they must logically be represented as producing a positive

¹ J. Chem. Soc., 1930, 443.

² Ibid., 1926, 1307: Annual Reports, 1926.

³ Solvay Reports, 1931; J. Chem. Soc., 1932, 1445; Institute of Chemistry Lectures, 1932.

field, $+\mathbf{F}$ instead of $-\mathbf{I}$, and conversely that basylous radicals (which include negatively-charged atoms) produce a negative field, $-\mathbf{F}$, instead of $+\mathbf{I}$. Ambiguity on this account may be avoided by using arrows to represent the direction in which electrons are displaced by the fields in question, since (fortunately) the arrows used to represent dipoles point in the same direction as those which indicate the drift of electrons produced by them, thus:

The reactions influenced by electrical fields have also been classified by Ingold and Rothstein in relation to the vector sign of the polar field. Those reactions which are facilitated by a displacement of electrons in the direction of a reacting atom or group are termed Class A reactions, and those which are facilitated by a corresponding displacement away from the reacting group Class B reactions: e.g.

Class A. Hydrolysis of Alkyl halides in neutral or acid solution.

Class B. Dissociation of a Carboxylic acid.

Cl—CH₃—CO·O—H
$$\rightarrow$$
 Cl—CH₃—CO·O + H
 \leftarrow + + F increases reactivity.
CH₃—CH₃—CO·O—H \rightarrow CH₃—CH₂—CO·O + H
+ \rightarrow - F decreases reactivity.

A warning that the classification as A or B applies to the reaction as a whole, must however be given, since it is never

¹ J. Chem. Soc., 1928, 1217.

possible to elucidate completely the polar mechanism of a reaction by examining (as above) the behaviour of only one of the reactants. Thus, in the Class A reaction instanced above—the hydrolysis of an alkyl halide—one has to remember that though a polar field $+\mathbf{F}$ will diminish the ease of ionization of the C—Cl bond, to C—Cl, it will attract the hydroxyl anions which interact with the activated molecule and complete the chemical change

 $R-C_6H_4-CH_5-Cl+OH\rightarrow R-C_6H_4-CH_5-OH+Cl$ In the particular example chosen the effect of the polar field (F) upon the hydroxyl ions (present in very low concentration) is less potent than its effect upon the C-Cl bond in controlling the rate of reaction. On the other hand, in alkaline solution the effect of the polar field upon the reactant anion predominates, as is the case also in the reaction between an alkyl halide and iodide anion for which data have already been given (p. 227).

$$\overset{-}{\overset{+}{\text{O}}} - \overset{+}{\overset{-}{\text{N}}} - \overset{-}{\overset{+}{\text{C}}} + \overset{-}{\overset{-}{\text{C}}} + \overset{-}{\overset{-}{\text{I}}} + \overset{-}{\overset{-}{\text{C}}} + \overset{-}{\overset{-}} + \overset{-}{\overset{-}{\text{C}}} + \overset{-}{\overset{-}} + \overset{-}{\overset{-}{\text{C}}} + \overset{-}{\overset{-}{\text{C}}} + \overset{-}{\overset{-}{\text{C}}} + \overset{-}{\overset{-}{\text{C}}} + \overset{-}{\overset{-}{\text{C}}} + \overset{-}{\overset{-}}} + \overset{-}{\overset{-}{\text{C}}} + \overset{-}{\overset{-}} + \overset{-}{\overset{-}{\text{C}}} + \overset{-}{\overset{-}}} + \overset{-}{\overset{-}{\text{C}}} + \overset{-}{\overset{-}}} + \overset{-}{\overset{-}}} + \overset{-}{\overset{-}} + \overset{-}{\overset{-}}} + \overset{-}{\overset{-}}} + \overset{-}{\overset{-}} + \overset{-}{\overset{-}}} + \overset{-}{\overset{-}} + \overset{-}{\overset{-}}} + \overset{-}{\overset{-}} + \overset{-}{\overset{-}}} + \overset{-}{\overset{-}}} + \overset{-}{\overset{-}}} + \overset{-}{\overset{-}} + \overset{-}{\overset{-}}} + \overset{-}{\overset{-}}} + \overset{-}{\overset{-}}} + \overset{-}{\overset{-}}} + \overset{-}{\overset{-}}} + \overset{-}{\overset{-}}} + \overset{-}} + \overset{-}{\overset{-}}} + \overset{-}{\overset{-}}} + \overset{-}{\overset{-}}} + \overset{-}{\overset{-}}} + \overset{-$$

 \leftarrow + + F attracts \overline{I} : Class B reaction.

The Origin of General Polarity.

The electric field required for the production of general polarity can originate in two ways.

(a) Positive and Negative Poles. Ingold 2 uses the term pole to describe the positively- or negatively-charged atoms which can generally be postulated in polyatomic ions (e.g. the nitrogen in the kation of an ammonium salt, or the oxygen in the anion of a phenate or carboxylate). Any such "bound ion" will produce, at a distance r, in a medium of dielectric constant D, a field of intensity, $E = e/Dr^2$, and an electrical potential $\phi = e/Dr$. The field E is of

¹ Franzen and Rosenberg, J. prakt. Chem., 1920, 101, 333

² Chemical Society Annual Reports, 1926, p. 129.

considerable magnitude for distances of the order of atomic dimensions. Thus Debye ¹ has pointed out that the field at the surface of an isolated univalent ion in water may be as much as 600 e.s.u., and the local potential gradient 180,000 volts per cm. In solvents of low dielectric constant, or across free space, the forces may be greater still. The change of energy, E_D , in a proton or electron, resulting from the "direct" action of an induced potential ϕ_D is $E_D = Ne\phi$ per gram-molecule. Values calculated for typical molecules are of the order of 100 to 10,000 calories per gram-molecule, ² and these are ample to account for induced changes of reaction-velocity, since the energy of activation of an organic reaction is usually of the order of 20,000 calories.

(b) Dipoles. Many molecules contain permanent electrical dipoles (Chapter IV) of the order of 10⁻¹⁸ e.s.u. per molecule. Nitro-compounds, sulphoxides and sulphones which contain semi-polar bonds generally have large permanent dipoles, whilst smaller moments are associated with covalent bonds between atoms of different atomic number.

A permanent dipole will produce, at a distance r, in vector orientation θ to its length, a field of intensity

$$E' = \frac{\mu\sqrt{(1+3\cos^2\theta)}}{Dr^3}$$

and 3 potential $\phi'=\mu\cos\theta/Dr^2$. This potential is of the order of a few tenths of a volt at molecular distances. Langmuir 4 has pointed out that the electrical forces between dipoles of magnitude 10^{-18} e.s.u. are negligible unless molecules are in actual contact with one another; and even then the work required to separate the molecules, as measured by their mutual heat of evaporation, is small.

^{1&}quot; Polar Molecules", p. 112.

² Waters, J. Chem. Soc., 1933, 1551.

³ Waters, Phil. Mag., 1929 (viii), 436; J. Chem. Soc., 1933, 1551; cf. also Smallwood, J.A.C.S., 1932, **54**, 3052; Schwarzenbach and Egli, Helv. Chim. Acta, 1934, **17**, 1183.

⁴ Langmuir, Chemical Reviews, 1929, 6, 451; cf. Lennard-Jones, Proc. Phys. Soc., 1931, 43, 461.

The electrical forces within molecules may, however, be very much greater, both on account of the lesser distance of separation and on account of favourable orientation of the linkages concerned. At a very short distance, therefore, the resultant electrical field of a dipole may be almost as much as that due to a single pole.

The Effect of General Polarity upon Reaction Velocity.

In a previous chapter it has been pointed out that a definite amount of energy of activation must be supplied before a covalent bond can be broken (p. 82). This energy of activation, E, is one factor in determining the reaction-velocity, k, since this can be expressed by the equation

$$k = Be^{-E/RT}.$$

In an earlier chapter (p. 83) it has been mentioned that Moelwyn-Hughes has shown that for the reactions between covalent bonds and ions the factor B is nearly equal to the frequency of collision of the reactants, and is independent of their structures.

For some typical slow reactions between two covalent molecules too it was shown by Bradfield 1 that the factor B is not altered by varying the nature of a "polar" substituent group in one of the molecules, provided it is not vicinal to the point of reactivity. From more recent investigations Williams and Hinshelwood 2 conclude that "whatever the nature of B, it remains of the same order of magnitude in any given solvent", and that "the very marked internal electronic displacements . . . caused by substituents, may change E considerably, while having a far less effect on B".

One can state, therefore, that General Polarity influences particularly the energy of activation of chemical reactions.

¹ J. Chem. Soc., 1928, 1006, 3073.

² Ibid., 1934, 1081. See, however, Baker and Nathan, ibid., 1935, 1840; 1936, 236; Ingold and Nathan, ibid., 1936, 222; Hughes, Ingold and Shapiro, ibid., 1936, 225; Pickles and Hinshelwood, ibid., 1936, 1353.

Its effect can be examined as follows:

The energy of activation E of a reaction can be resolved into two parts, namely: (i) a constant E_o representing the intrinsic energy of activation, under given experimental conditions, of the bond which is being ionized, e.g. of the C—Cl link in a series of alkyl chlorides undergoing hydrolysis, and (ii) a variable ΣE_p representing variations in this energy induced by general polarization from more distant atoms. Thus

$E = E_o + \Sigma E_p.$ (a) Influence of Internal Fields.

Many organic reactions in solution have been shown to be unimolecular processes, and all these therefore have velocities depending upon the speed of activation of only one organic reagent. When considering the relative reactivities of an isotypic series 2 of similarly constituted compounds towards one particular reagent, bringing about an ionic reaction, under conditions which have been standardised in the experimental investigation, one can, in these cases, take E_0 as the constant energy of activation of the parent unsubstituted compound and ΣE_p as the influence of the extra electrical fields in the substituted molecules upon the ease of ionization of the link which breaks. Then, if one finds experimentally that the factor B is constant throughout the series, as is often the case, one can deduce from the general expression

$$k = Be^{(E_o + \Sigma E_p)/RT},$$

that

 $Log rac{k_{subs.~cpd.}}{k_{unsubs.~cpd.}} = const. imes \Sigma E_p$. . . a measure of the polar effects.

The relative rates of reaction of several series of compounds have been shown to correspond approximately with values calculated for ΣE_p according to the electrostatic conception given in the previous section, thus justifying the theoretical basis of the electrical theory of general polarization. Good examples are afforded by the changes in the dissociation constants of the dibasic organic acids

¹ Hughes and Ingold, J. Chem. Soc., 1933, 526, 1571; 1935, 236. 244; 1936, 225.

⁸ Ingold, J. Chem. Soc., 1930, 1035. Cf. page 278.

(Chapter IX) and the rates of hydrolysis of their esters (Chapter XII) where the polar action is due to the anion

—CO·O operating over varying distances. Dipolar groups also produce effects upon reaction velocity in proportion to the magnitude of their dipole moments. Thus the strengths of *meta*-substituted aromatic acids follow the magnitudes of the dipole moments of the substituents, and for acids of the general type X·CH₂—COOH the dissociation constant obeys very closely an equation of the type

RT log $k = \text{const.} + c(\mu + \alpha \mu^2)$ where $\alpha = 0.032$, as can be seen from the following table:

Table I
Values for log K for X—CH₂—COOH

	Log K					Log K	
X	μ	found calc.	х	μ	found	calc.	
CH_3		II·2	Br.	- 1·8o	6.59	- 6.83	
Η.	0	- 10·9	Cl .	- I·88	-6.47	- 6.66	
Ph.	o·39	- 9·8 - 10·0	Cl ₂ .	- 3.76	- 2.97	- 2.96	
MeO	- I·29	– 8·o – 7·9	Cl ₃ .	- 5.64	+ 0.19	+ 0.17	
Ι.	- 1.63	-7.19 - 7.19					

Nathan and Watson,² to whom the above table is due, have found that a similar correlation between relative reaction velocities and dipole moments can be deduced for other reactions in solution between ions and molecules needing activation. It is remarkably exact when changes in collision frequency (i.e. in B) are taken into account. For the slow reaction between primary aromatic amines and acyl chlorides in benzene solution a similar functional relationship between k and μ has been disclosed by Williams and Hinshelwood.³

(b) Influence of one Reactant on the Other.

The equations of (pp. 244-247) can be used to calculate the energy change E_p due to a substituent group not only

¹ Waters, Phil. Mag., 1929, viii, 436.

² J. Chem. Soc., 1933, 890, 1248.

³ Ibid., 1934, 1082.

in one reagent molecule, but also if the polar group is in the other reactant. Thus if one is considering an additionreaction, indicated by

(1)
$$\overrightarrow{R} - \overrightarrow{A} = \overrightarrow{B}$$
 (see p. 241) $\overrightarrow{X} - \overrightarrow{Y}$

the substituent group R will influence the ease of ionization of both the bonds A-B and X=Y, though undoubtedly its influence will be very much less than the mutual inductive effect of the two reactants, discussed on page 240. The resultant effect of a substituent R upon the two reactant groups is very complex, since it depends in great part upon their relative tendencies towards activation, and if, as in diagram I above, the field due to the substituent R promotes activation of one of the reactant groups (A=B) it will tend to inhibit activation of the other. Consequently one cannot directly correlate electrostatic field strength (F) with reaction velocity when there are two variable reactants, and it is even possible for the general polar class (A or B, p. 243) of a reaction to change with change in magnitude but not sign of F. Such a case is indicated in Table II.

Table II

Reaction of $p \cdot Y \cdot C_0 H_4 - CH_2 - Br$ with aniline in 90% Ethyl

Alcohol 1 $Y = CH_2 - H$ I Br Cl NO

Such anomalous results as those of the above table have been ascribed to "reversed field effects", but it is not necessary that the vector direction of \mathbf{F} should change whenever the resultant polar effect, ΣE_p , changes sign.

In the simpler case of a reaction between a dipole and an ion, Spong 3 has pointed out that since an ion is repelled

¹ Baker, J. Chem. Soc., 1933, 1128.

² Baddeley and Bennett, J. Chem. Soc., 1933, 261; Bennett and Mosses, J. Chem. Soc., 1930, 2364.

³ J. Chem. Soc., 1934, 1289.

from one end of a dipole but attracted by the other, its frequency of collision with the dipole will not be arithmetically calculable from a knowledge of its bulk concentration.



If n is the bulk concentration of the ion, then in a field of potential V its local concentration $n' = ne^{\frac{\epsilon V}{RT}}$, where ϵ is the ionic change, k is Boltzmann's constant, and V is approximately equal to $\mu \cos \theta/Dr^2$.

Hence if Z is the collision number calculated from the bulk concentration of the reactants (i.e. from n and not n')

$$k = Ze^{\frac{-(E-NoeV)}{RT}}$$

where N_o is Avogadro's number and E the true energy of activation of the reaction.

For most reactions it is not possible for the reactant ion to approach the dipole with equal facility in all directions, and, further, the requisite distance of approach r of the ion to the dipole before reaction ensues is again a variable function depending upon the orientation of the molecule (i.e. on θ). Hence the mean value of V is not zero. $(N_o eV)$ will represent a positive energy, to be added to the energy of activation E if the ion is, on the average, repelled by the dipole, and vice versa. In this way it is possible to discover the orientation on a complex molecule, such as sulphur chloride, of alocal dipole capable of reacting with an ion.

(c) Influence of External Fields. In reaction mixtures, especially in ionizing solvents, free ions can affect the electrical environment both of ionized reagents and of the electrons which form the vulnerable bond of the reactant. The phenomenon of electrostriction, which is concerned with the effect of free ions upon each other, has been discussed quantitatively by Milner, Debye and

Hückel, Onsager and others 1 in their theories of strong electrolytes. The principal effect of this interaction is to diminish the mobility of the ions and thus to decrease the apparent "degree of dissociation" even of electrolytes which are completely ionized in solution. The equations which represent the effect of electrostriction indicate that the difference in energy of an ion of valency z when isolated, and when in presence of a concentration c of ions in a medium of dielectric constant, D, is proportional to

$$\frac{z^2\sqrt{\Sigma(cz^2)}}{D^{3/2}T^4}$$

where Σcz^2 represents the sum $c_1z_1^2 + c_2z_2^2 + \ldots$ for each ion present. This expression, however, is only valid for ionic concentrations not exceeding 0.001N or thereabouts.

Fajans ² has stressed the fact that neighbouring ions in crystal lattices, as well as in solution, can deform each other, producing changes of colour, refractivity and shape. In this way the difference in colour between the crystals and the dilute aqueous solutions of mercuric and lead iodides can be explained. Moreover, ions and polar molecules can associate together even in dilute solutions to form ion-pairs, or still more complex aggregates. These aggregates will have smaller external fields, but may have some stronger internal fields than the isolated ions.

Since the ionic reactions of covalent molecules depend ultimately on interaction between ions, it might be expected that reaction velocities would be influenced by the same factors. Thus (i) The Debye-Hückel-Onsager equation would suggest that the energy-level of a reactant molecule in solution would be related in some simple way to the dielectric constant of the solvent; but it by no means follows that reaction velocity should be capable of simple correlation with the dielectric constant, since both the factors B and E_o of the reactivity equation can alter profoundly with change of solvent, and in fact no consistent way of

¹ Cf. Chemical Society Annual Reports, 1930, pp. 326-56.

² G. F. Baker Lectures, Cornell University, 1930.

correlating reaction velocity with dielectric constant has vet been found.1 The bulk dielectric constant of the solvent medium may not be a valid measurement for particles at molecular distances from one another. Further, a dissolved substance often associates with the solvent to form "solvates" which may differ widely in electronic constitution in different media.

(ii) Robinson 2 has reported instances in which the relative rates of two competing reactions are affected profoundly by the presence of other ions in the reaction mixture. For instance, aromatic bases are nitrated in the ortho- and parapositions when weak acids are used, and in the meta-position by strong acids, but added salts decrease the percentage of meta-isomeride. Thus benzyl-piperidine, when nitrated with strong nitric acid at oo, gave 72% of meta-isomeride; but the percentage of meta-isomer was reduced to about 55% by the addition of large amounts of ammonium nitrate, to 25% by rubidium nitrate, and to under 20% by trimethylammonium nitrate. Addition of organic molecules, such as nitrobenzene, produced a much smaller effect. Robinson suggests that the effect of electrostriction should increase with the magnitude of the external inductive field, and with it the mobility (polarizability) of the groups which are exposed to the influence of this field.

The Transmission of Polarity Within Molecules.

The way in which the polarization, or electrical dissymmetry of a radical in one part of a chemical molecule produces its general polar effect at more distant parts of the same molecule has been a matter of much theoretical discussion. It is now agreed that the "polar effect" of a radical is exerted partly in virtue of an electrical field which acts directly across space, but may also be transmitted by a process of electrostatic induction from atom to atom, or

¹ Cf. Moelwyn-Hughes, "Kinetics of Reactions in Solution", 1933, p. 118.

² J. Chem. Soc., 1927, 2770.

³ Compare page 446.

bond to bond, within the molecule. Obviously a dipolar radical can polarize an adjacent covalent bond to produce in it a small secondary dipole, which again can induce electrical polarization in the next bond, just as a bar magnet can induce magnetization in a chain of soft iron filings.

Lapworth 1 has suggested that this internal induced effect, which may represent an exceedingly efficient mode of transmission of polarity, should be termed the *imported effect*, to distinguish it from the primary "direct" or "field" 3 effect which is exerted directly across space. Ingold 2 confines the term "induced effect" to this internally transmitted polarization, denoting it by the symbol I, but this is apt to cause confusion since it has so long been customary to speak of all manifestations of general polarity as "induced polarity".

Whilst the magnitude of the "direct" or "field" effect of a polar or a dipolar group upon chemical reactivity can be calculated approximately on the lines indicated in the previous sections, any theoretical estimate of the magnitude of the internally imported effect is most uncertain, as it necessitates a very precise knowledge of molecular structure.

One can approach the problem, however, as follows: 4

If α is the polarizability of a bond, defined so that in a field F the bond acquires an induced electrostatic moment $m=\alpha F$, then, in the field F the bond will gain electrostatic energy $\Delta E=\frac{1}{2}\alpha F^2$. Since the field F due to a pole at distance d is e/d^2 , and due to a dipole varies as μ/d^3 the induced energy in a polarized bond will vary with e^2 in the field due to a pole, and with μ^2 in a field due to a dipole.

The magnitude of the dipole moment induced in a chain of atoms will diminish in geometrical progression as the number of valency linkages increases, and will very soon become negligible unless the polarizability of each link is high. Now the polarizability of a bond can be measured

¹ J. Chem. Soc., 1930, 445. ² Ibid., 1928, 418, 2262.

³ Ibid., 1932, 1445; Solvay Report, 1931, etc.

⁴ Waters, J. Chem. Soc., 1933, 1554; cf. Ingold, Chemical Reviews, 1934, 15, 226.

by its refractivity (see pp. 56-58, Chapter IV) and, by substituting bond refractivities in the equation $E = \frac{1}{4}\alpha F^2$. one finds that, in a saturated carbon chain each bond should acquire by induction about 0.02 of the energy of the previous one, whilst an olefinic link could acquire as much as 0.5 of the energy of the preceding link. Hence the induction of polarity should be negligible after traversing about three bonds in a saturated chain, whilst an unsaturated linkage should produce an effect on polar activation equivalent to that of a conductor placed in a polarized dielectric. If a bond acquires electrostatic energy by induction, then its electrons must rearrange themselves from their original, to new "perturbed" orbits, and undoubtedly this perturbation will take place so that the electrons move from higher to lower states of potential energy in the imposed electrostatic field. Hence the "induced" or "imported" effect will always be represented by a term $+\hat{E}_{I}$ where +Erepresents energy that must be gained before activation can occur.1

One can thus account for the existence of the square term, $+c\mu^2$ in the empirical equation by which Nathan and Watson successfully correlated the activation energy of a chemical change with the dipole moment of the substituent group present (p. 248).

This "square term" will also appear in any empirical equation, that can be used to correlate, with activation energy, the induction of polarity in different reacting molecules.

Verification of the above conception of the nature of "induced" polarity can be obtained in two ways since—

- (i) By accurate examination of the dipole moments of complex molecules it is possible to prove the existence of induced dipoles and estimate both their magnitudes and vector orientations.
- (ii) By examination of refractive indices one can compare the polarizabilities of different covalent bonds. Correlation of bond polarizability with chemical reactivity then reveals

¹ Waters, Nature, 1934, 134, 178.

the fact that the greater the polarizability of a linkage the more susceptible it is to general polar influences.

Experimental evidence from both these lines of approach is discussed below.

Induced Dipole Moments.

By comparing the dipole moments of radicals such as OH and NO₂ in different types of compounds, Sutton ¹ has shown that the measured dipole moment is not absolutely constant from compound to compound, but exhibits a slight variation which can be correlated with molecular structure, and which indicates that each molecule possesses not only its primary dipole but has in addition a small internally induced dipole.

TABLE III ^a

Moments of Substituent Groups (in 10⁻¹⁸ e.s.u.)

						D:0	·
	Aromatic Seri	es.		Aliphatic Serie	es.		ierence -Alk).
Aryl	CH_3	+ 0.45	Alkyl	⊢CH₃	0.00	-1-	0.45
,,	OCH ₃	— 1.00	,,	OCH ₃	- 1.29	+	0.23
,,	$-NH_2$	+-1.55	,,	-NH ₂	+ 1.23	+	0.32
,,	Cl	- 1.56	,,	Cl	- 2.15	+	0.59
,,	Br	- I·52	,,	Br	- 2·2I	+-	0.69
,,	—-I	- 1.27	,,	I	- 2.13	+	o·88
,,	CH ₂ Cl	-1.82	,,	CH₂Cl	- 2.03	+	0.51
,,	CHCl2	- 2.03	,,	CHCl ₂	- 2.06	+	0.03
,,	CCl ₃	- 2.07	,,	—CCl ₃	- 1.57		0.50
,,	COCH ₃	2.97	,,	-COCH ₃	- 2.79		0.18
,,	-co	- 3.04	,,	-=CO	- 2.76		0.28
,,	CN	- 3.89	,,	CN	- 3.46		0.43
,,	NO_2	- 3.93	,,	$-NO_2$	- 3.05		o·88
	-			_			

If there were no electrical induction between the groups, we should expect to find a constant difference between $C_0H_0\bar{X}$ and CH_3X , which could be accounted for by assigning a fixed dipole moment of + 0.45 to the phenyl radical, just as we assign a fixed dipole moment of - 2.15 to the chlorine radical. The table shows, however, that the difference between the values for the fatty and aromatic series is not only extremely variable, but actually under-

¹ Proc. Roy. Soc., 1931, A 133, 668.

² From Sidgwick, The Covalent Link in Chemistry, 1933, p. 179.

goes a reversal of sign. This effect can only be explained by postulating a mutual induction between the radicals.1 We might suppose, for instance, that the substituent X has a constant moment and that the numbers in the last column represent the dipole moment of the phenyl group, when modified by the inductive action of the more powerful moment of the substituent X, the values of which are tabulated in the second column; but an effect of this kind is almost certain to be reciprocal and it would be unwise to attribute the irregular differences in the last column exclusively to the polarization of the phenyl radical. data do, however, establish beyond question the fact that dipoles in a molecule may exert a mutual induction upon one another, and that the transmission of polarity along a chain of atoms is therefore possible, at least in the case of those atoms or radicals which are most readily polarized.

Bond Polarizability.

It has already been shown in Chapter IV that the *total* polarization, P, of any molecule, as measured by its dielectric constant, is made up of three parts:

$$P = P_{\mu} + P_{A} + P_{E}$$

where P_{μ} depends on the orientation of permanent dipoles, P_{A} depends on the displacement of positive nuclei, P_{E} depends on the displacement of electrons.

Light which is of too high a frequency to effect a displacement of nuclei can only produce a displacement of electrons, so that the molecular refraction of visible light is a measure of P_E only.

$$P_E = \frac{4}{3} \pi N \alpha_0 = \frac{n^2 - 1}{n^2 + 2} \cdot \frac{M}{d}$$

where N is the Avogadro constant;

 α_o is the molecular polarizability, as measured by the ratio between the electric field and the dipole moment induced by it;

n is the refractive index, which can if desired be extrapolated to infinite wave-length (from

¹ Molecular stabilization by *resonance* has also to be considered. Compare page 395.

observations which are not influenced by infrared absorption bands) in order to eliminate the effects of dispersion.

Molecular refractivity has long been regarded as an additive property, which can be resolved into a series of atomic refractions, with suitable corrections for various types of binding. Since, however, it is the electrons and not the atoms that are concerned in the refraction of light, it is more logical to assign constant values to the various electron-groups, rather than to the atoms themselves. This method of analysis, due originally to von Steiger, has been developed both by C. P. Smyth 2 and by Fajans and Knorr, all of whom have calculated BOND REFRACTIVITIES. From these bond-refractivities the various BOND POLARIZABILITIES can be calculated, by using the same formula as for molecular refractivities.

TABLE IV

Refractions of Electron-groups, P_B.

H—H	2.08	CH	1.705	C-C	1.209
COH	3.76	C-Cl	6.57	C = C	4.15
COC	2.85	C-Br	9.47	c = c	6.025
C=O	3.42	C—I	14.51	C=S	10.61
C-NH,	5.13	C ₂ :NH	4.81	C_3N	4.65
(Primary am	ine.)	(Secondar	y amine.)	(Tertia	ry amine)

The following deductions can be drawn from these data:

(i) The Halogens. The polarizability of the halogens increases rapidly in the series F < Cl < Br < I. This is true not only of the carbon-halogen link but also of the halide ions and the halogen hydrides (Fig. V). Thus the outermost electrons of the large atoms of iodine are more easily displaced by electric fields than are the electrons of fluorine, which are much nearer the central nucleus. These results are concordant with the sequence of reactivity, both for non-polar dissociation, as in the thermal or photo-

¹ Ber., 1921, 54, 1381.

^{2&}quot; Dielectric Constants and Molecular Structure", 1931, p. 152.

³ Ber., 1926, 59, 256.

chemical decomposition of the halogen acids, and for polar reactivity as exemplified by hydrolysis, reduction and substitution (pp. 135 and 225).

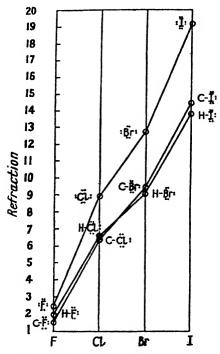


FIGURE V.1—Refractions of the Electron Groups of the Halogens, from C. P. Smyth's "Dielectric Constants and Molecular Structure", p. 156.

(ii) Multiple Linkages. The refractivity of a single bond between carbon atoms is 1.209 units, for a double bond it is 4.151 units, or 2.076 per electron-pair, whilst for a triple bond it is 6.025 units or 2.01 per electron-pair. This indicates that the double bond is almost twice as easily polarized as a single bond, but differs very little

¹ Reproduced, from C. P. Smyth's, "Dielectric Constants and Molecular Structure," 1931 (By permission of the Rheinhold Publishing Corporation, New York).

from a triple bond in its polarizability. Conjugated systems of multiple linkages are slightly more polarizable than isolated double bonds, but for aromatic rings the total polarizability is slightly less, indicating that the formation of an aromatic ring is conducive to stability. The polarizability of multiple linkages provides, however, a simple explanation of the way in which they can transmit the effects of "general polarity" through distances which may be much greater than those which could be reached by an unaided field of electrostatic force.

Thus in Chapter IX it has been shown that there is a close similarity in the variations of the dissociation constants of the carboxylic acids in the series R·COOH and R·CH:CH·COOH as the growing chain R is extended. This can now be interpreted in terms of general polarity by regarding the unsaturated group —CH:CH— as a system which is very easily polarized. If, therefore, the radical R carries a dipole, which exerts a "direct" or "field" effect on the carboxyl group in R·COOH, we may expect that the olefinic group in R·CH:CH·COOH will develop an induced dipole moment, the effect of which will be very similar to that of bringing R into direct relation with the carboxyl group, e.g.

R·CO·OH R·CH:CH·CO·OH

where the induced dipole is indicated by a broken arrow.

(iii) Conjugated and Aromatic Systems. Conjugated chains of alternate single and double bonds differ in toto from saturated aliphatic chains, in that they can exhibit an alternation of polarities, the mechanism of which will be discussed in a subsequent chapter (Chapter XV). For the moment, therefore, it will suffice merely to record the fact that positive poles tend to induce meta-substitution in an aromatic nucleus, e.g. aniline gives ortho- and paranitro-derivatives, but the trimethylammonium ion gives a meta-nitro-derivative. The following observations can then be recorded to illustrate the rapidity with which polarity decays in a saturated system.

260 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

(i) Ingold 1 has shown that in the series of compounds

$$C_6H_5\cdot[CH_2]_n\cdot NMe_8$$

the percentage of meta-nitration falls off as follows:

(ii) The carboxyl group in benzoic acid is also metadirecting; but phenylacetic acid yields o-p-derivatives, like toluene:

.(iii) The nitro-group is also *meta*-directing, so that the nitration of nitrobenzene yields almost exclusively m-dinitrobenzene. Similarly phenylnitromethane yields mainly *meta*-derivatives, but ω -nitrostyrene gives mainly o-p-derivatives.²

In the last case, the olefinic group, instead of transmitting the polarity of the substituent as it does in the unsaturated fatty acids, appears to impose its own orienting influence on the ring.

(iv) The three nitro-diphenyls also yield o-p-derivatives in the ring which does not contain the nitro-group, showing that the polarity of the nitro-group is not transmitted from one aromatic ring to the other (cf. p. 466).

In general, we conclude that substitution in aromatic nuclei may be governed by positive or negative polarities transmitted from the substituents or side-chains. These set up a system of alternating polarities in the ring, which leads either to m- or to o-p-substitution (cf. Chapter XVI); but the effect dies out very rapidly as the substituent in

¹ J. Chem. Soc., 1927, 250.

² Baker and Ingold, *J. Chem. Soc.*, 1926, 2462; Baker and Wilson, *J. Chem. Soc.*, 1927, 842.

the side-chain becomes more remote; and it is easily blocked or reversed by the interposition of groups which have definite polarities of their own. On the whole, these effects are what one might expect in the complex aggregate of nuclei and electrons which makes up a molecule, since there are obvious electrostatic analogies for the transmission of a field either directly through space or through polarizable particles, as well as for the shielding effects which are sometimes indicated.

CHAPTER XII

HYDROLYSIS AND ESTERIFICATION

Hydrolysis an Ionic Reaction

The term "hydrolysis" is commonly used to describe any chemical change in which water is one of the reactant molecules. The present chapter deals, however, only with those changes in which a covalent bond is broken with addition of hydrogen and hydroxyl radicals to the two radicals which were initially united by the bond. Hydrolyses of this type are particularly common amongst the derivatives of carboxylic acids:

(a) Esters, lactones, etc.

$$R \cdot CO \cdot OR' + H_2O \longrightarrow R \cdot CO \cdot OH + HO \cdot R'$$

- (b) Amides $R \cdot CO \cdot NH_2 + H_2O \longrightarrow R \cdot CO \cdot OH + NH_3$
- (c) Acyl chlorides $R \cdot CO \cdot Cl + H_2O \longrightarrow R \cdot CO \cdot OH + HCl$.

These hydrolyses proceed slowly and with a well-defined energy of activation. They may, therefore, be differentiated sharply from the hydrolysis of salts, e.g.

$$CH_3 \cdot COO \overset{+}{Na} + H_2O \rightleftharpoons CH_3 \cdot COOH + \overset{+}{Na} \overset{-}{OH}$$

which is a mere reversal of the process of neutralization, and therefore takes place instantaneously and without requiring any energy of activation (p. 83).

Without exception, hydrolyses are ionic reactions, since the water acts invariably as a source of hydrogen and hydroxyl ions, instead of being resolved into neutral hydrogen and hydroxyl radicals. Reactive substances, such as acetyl chloride, can be hydrolysed by pure water, which contains sufficient hydrogen and hydroxyl ions to initiate the change. In general, however, the hydrolysis must be promoted by the addition of a catalyst in order that the chemical change may take place with a reasonable velocity. These catalysts are substances of polar nature, such as the mineral acids, the caustic alkalis, or certain enzymes of protein (i.e. of "zwitterion") nature, all of which are capable of supplying either hydrogen or hydroxyl ions to the reaction mixture. Hydrolyses therefore depend on the interaction of the hydrolyte with donors of hydrogen and hydroxyl ions (cf. pp. 312-316, Chapter XIII), and can be investigated very conveniently from the standpoint of theories of general polarity, since the velocity of hydrolysis may be influenced profoundly by the extent to which these ions are attracted or repelled by the dipolar molecule.

In order to simplify the treatment of the wide subject of hydrolysis, only the hydrolysis of esters and lactones of carboxylic acids will be considered in detail. This reaction depends on the rupture of a covalent linkage between carbon and oxygen by the addition of the radicals of water; but it differs very little in its manner of occurrence and essential mechanism from the rupture of a carbon-nitrogen linkage in the amides, or of a carbon-halogen linkage in the acyl halides.¹

The Mechanism of Hydrolysis and of Esterification.

(a) Solvents. The hydrolysis of esters is most conveniently effected in solution. It can also take place, however, though much less rapidly, at the interface of two different phases, as in the saponification by an alkali of a fat which is insoluble in water. The utility of the Twitchell reagent as an emulsifier in promoting the hydrolysis of fats is, however, an indication that an intimate mixing of the ester with the aqueous catalyst is an important factor in facilitating the change. Solvents such as the alcohols may be used to provide a medium in which all the reactants are soluble; but solvents which can provide hydrogen or

¹ Cf. Reid, Amer. Chem. J., 1899, 21, 284; 1900, 24, 397; 1909, 41, 483; Berger and Olivier, Rec. trav. chim., 1927, 46, 516; cf. p. 235.

hydroxyl ions may also act in a dual capacity as solvents and as catalysts for the hydrolysis.

- (b) Catalysts. Hydrolysis of an ester is usually effected in either acid or alkaline solution. The second case is often described as saponification, but does not differ essentially from acid hydrolysis.
- (i) Hydrolysis by Acids. The equation for the hydrolysis of an ester by an acid includes a molecule of water, but the acid itself need not be shown in the equation, except as a catalyst, since it is not used up in the process:

$$CH_3 \cdot CO \cdot OC_2H_5 + H_2O \xrightarrow{HCI} CH_3 \cdot CO \cdot OH + HOC_2H_5$$

On the other hand, the carboxylic acid, which is formed as a product of hydrolysis, usually acts as an autocatalyst. It may accelerate the process very considerably, especially when the hydrolysis is carried out either in pure water, or in a medium containing only a small quantity of a mineral acid. The kinetics of hydrolysis are, therefore, more complex than those of an isomeric change (Chapter XIII), in which the composition of the medium remains constant throughout. The kinetics are also complicated by the fact that the chemical change is reversible, so that, as the action proceeds, hydrolysis is accompanied by the converse process of esterification. This is also catalysed by acids, and to an approximately equal extent, since the equilibriummixture which finally results (as in the case of ethyl acetate, acetic acid, ethyl alcohol and water, or the still simpler case of a lactone in equilibrium with a hydroxy-acid) is not affected substantially by the presence of the acid.

(ii) Hydrolysis by Alkalis. In striking contrast to hydrolysis by acids, the equation for the saponification of an ester in alkaline solutions includes a molecule of alkali, but water can be omitted from the equation, or shown only as a catalyst, e.g.

$$CH_3 \cdot CO \cdot O \cdot C_2H_5 + NaOH \xrightarrow{H_1O} CH_3 \cdot CO \cdot ONa + HOC_2H_5$$

It is indeed possible to saponify an ester by the action of potassium ethoxide in absolute alcohol, in spite of the fact

that the action can only be represented by an equation which shows the formation of a molecule of *ether*, e.g.

$$C_eH_s\cdot CO\cdot OC_2H_s + \overset{+-}{KO}C_2H_s \longrightarrow C_eH_s\cdot CO\cdot \overset{-+}{OK} + C_2H_s\cdot O\cdot C_2H_s$$

The effect of this drastic dehydration of the system is to reduce the velocity of saponification to 1/6,000 of the velocity in water. When saponification is effected with potassium hydroxide in absolute alcohol the velocity of saponification is reduced to 1/2,000 of the velocity in water 1; but the change cannot be arrested completely, since it is obviously impossible to deprive the alkali of all trace of water in a system which has been shown to give up water to potassium ethoxide.

Alcohol is, of course, a substance of the "water type", which can take part in acid-base reactions as an amphoteric compound. There is, therefore, ample justification for maintaining the view that, in the case of "alcoholysis" quoted above, the ethoxyl radical of the alcohol plays exactly the same part as a hydroxyl radical plays in aqueous saponification. In fact, any solution of an alkali hydroxide, even in aqueous alcohol, contains appreciable quantities of metallic alkoxide,²

$$NaOH + C_2H_5 \cdot OH \rightleftharpoons Na \cdot O \cdot C_2H_5 + H_2O$$

since reactions such as Williamson's synthesis of ether can be carried out in presence of a considerable proportion of water, and alcoholysis of esters can be effected in aqueous alcohol, e.g.

$$C_0H_5\cdot CO\cdot O\cdot C_0H_5 + \stackrel{+}{NaO}C_2H_5 \longrightarrow C_0H_5\cdot CO\cdot O\cdot C_2H_5 + \stackrel{+}{NaO}C_0H_5$$

(c) Formation of Addition-Compounds by Condensation of Ions. Since a catalyst may be regarded as "a reagent which is reproduced as a product of the action", we can suppose that the decomposition of the ester depends

¹ Cashmore, McCombie and Scarborough, J. Chem. Soc., 1921, 119, 970.

² Lobry de Bruyn, Rec. trav. chim., 1899, 18, 41.

⁸ Gibby and Waters, J. Chem. Soc., 1932, 2643.

⁴ Lowry, J. Chem. Soc., 1925, 127, 1375.

on bringing the catalyst into combination with the two reacting substances, and that this involves in each case the formation of a tertiary complex thus:

> Hydrolysis [Ester, H₂O, HCl] ¹ Saponification [Ester, NaOH, H₂O] ¹

Von Pechmann² actually isolated a binary addition-product of ethyl benzoate with sodium methoxide and represented it as an "atomic" compound formed by adding the two radicals of the reagent to the double bond of the ester, thus:

On the other hand, Maass and McIntosh,³ by working at low temperatures, were able to isolate the compound $CH_3 \cdot CO_2Et$, HBr, melting at -36° . They found that addition compounds of this class were salt-like substances, which were formed with a large evolution of heat; they also acted as electrolytes when allowed to melt, either with or without an excess of either constituent. In this case there can be little doubt that addition takes place at the oxidic oxygen, just as in the case of methyl ether,

$$CH_3 \cdot O \cdot CH_3 + HCl \rightleftharpoons (CH_3)_2 \overset{+}{O}H + \overset{-}{Cl}.$$

Further evidence for compound-formation between esters and acids has been obtained by Kendall and his collaborators 4 from the analysis of freezing-point data. They also consider that addition-compounds can be formed between esters and water. It is now generally admitted that esters are "amphoteric" compounds, possessing both acidic and basic properties, and that the catalytic action of acids and bases is associated with this ability to combine with the ester, rather than on the emanation of some nebulous catalytic influence.

These binary addition-products must, however, be formu-

¹ The second term in each bracket represents the reagent which is used up in decomposing the ester, whilst the last term represents the catalyst.

² Ber., 1898, **31**, 503. ³ J.A.C.S., 1912, **34**, 1272.

⁴ J. Chem. Soc., 1925, 127, 1768, 1778.

lated as salts, i.e. as made up of pairs of ions held together by electrostatic attraction and not by bonds, since any process of electron-sharing would at once create a surplus of electrons. Thus, in the case of methyl acetate being decomposed by aqueous sodium hydroxide or aqueous hydrogen chloride, the initial products of the action would be formulated ¹ as

It will be noticed that in each case a single ion of the alkali or acid has been added to the ester, the other ion retaining its original form; and that in each case a condensation has been effected by transferring the electric charge of the ion to another atom and thus forming a bond between the ester and the electrically-neutral radical. The addition of OH is, in fact, exactly parallel to the addition of CN in Lapworth's cyanhydrin reaction, whilst the addition of H

If the free ions are omitted, the complexes formed by the union of the ester with the alkali or acid may be written

is exactly like the formation of a substituted ammonium

ion by the addition of CH₃ to a nitrogenous base.

In neither case, however, has any part been assigned to the water, which acts as a catalyst in the saponification and as a reagent in the hydrolysis. It is, indeed, clear that hydrolysis requires the addition of both of the ions of water to the ester, whereas we have hitherto only added one of these ions. We may therefore suppose that a second ion,

¹ See Lowry, J. Chem. Soc., 1925, 127, 1380.

to supplement the one derived from the acid or alkali, is obtained from the water and that in the formation of a ternary complex this ion is fixed by a precisely similar mechanism to that described above. We can, therefore, formulate the two ternary complexes as

If we again omit the free ions, it will be seen that in these complexes we are dealing with two salts of the same bipolar molecule, namely

and that this has been formed by a twofold use of Lapworth's cyanhydrin mechanism, once with a negative and once with a positive ion. In order to complete the decomposition of the ester it is then only necessary to break the central bond in the usual way, by converting the covalency into an electrovalency, in order to neutralize all the electric charges and thus to leave the acid and alcohol as the final products of the hydrolysis.

Hydrolysis either by acids or by alkalis can then be represented by the following scheme:

$$CH_{3} \stackrel{\circ}{C} -O \cdot CH_{3} + H \stackrel{-+}{OH} \rightleftharpoons CH_{3} \stackrel{\circ}{C} - O \cdot CH_{3} \rightleftharpoons$$

$$CH_{3} \stackrel{\circ}{C} + O \cdot CH_{3} = CH_{3} \stackrel{\circ}{C} + HO \cdot CH_{3} \rightleftharpoons$$

$$OH \qquad H \qquad OH$$

¹ Compare Biilmann's "zwitterion", Annalen, 1912, 388, 330.

in which the alkyl group of the ester is eliminated as a neutral molecule of an alcohol and not as a free kation.

This scheme is in accord with the observations of Reid,1 who showed that thio-acids esterify with elimination of hydrogen sulphide and not water,

i.e.
$$C_6H_5 \cdot CO \cdot SH + H \cdot O \cdot C_2H_5 \rightarrow C_6H_5 \cdot CO \cdot O \cdot C_2H_5 + H_2S$$

not $C_6H_5 \cdot CO \cdot S \cdot H + HO \cdot C_2H_5 \rightarrow C_6H_5 \cdot CO \cdot S \cdot C_2H_5 + H_2O$

Moreover, C. K. and E. H. Ingold ² have shown that if an ester containing an alkyl radical capable of undergoing isomeric change when in the form of a free kation (cf. p. 319)

$$\left(\begin{array}{ccc} \overset{+}{\text{CH}} & \overset{+}{\text{CH}} & \xrightarrow{\text{CH}} & \xrightarrow{\text{C$$

is prepared, then no isomerization occurs during its hydrolysis by acids. The alkyl radical (R) cannot, therefore, be liberated as in (A)

(A)
$$X \cdot CO \cdot O \cdot R + H^+ \rightarrow X \cdot CO \cdot O \cdot H + R^+$$
,

but must separate as part of an alcohol molecule (B).

(B)
$$X \cdot CO \cdot O \cdot R + H + OH \longrightarrow X \cdot C \longrightarrow OH H$$

Since, however, hydrogen and hydroxyl ions cannot co-exist in substantial concentrations in the same solution, we must suppose that the negative pole in acid solutions, and the positive pole in alkaline solutions is derived from a water *molecule*, which gives up a hydroxyl ion to the ester in acid solutions and a hydrogen ion to the ester in

¹ Amer. Chem. J., 1910, **43**, 489; J.A.C.S., 1915, **37**, 1934; 1916, **38**, 2746.

² J. Chem. Soc., 1932, 756. Compare also Polanyi and Szabo. Trans. Faraday Soc., 1934, 30, 512.

⁸ Lowry and G. F. Smith, J. Chem. Soc., 1927, 2539.

alkaline solutions. The hydrolysis is thus represented by a scheme of one of the following:

Acid Hydrolysis: $\overset{+}{O}H_3 + X \cdot Y + 2OH_2 \longrightarrow OH_2 + HX + YOH + \overset{+}{O}H_3$ Alkaline Hydrolysis: $OH_2 + X \cdot Y + OH \longrightarrow OH + HX + YOH$

The positive pole in this process can be provided by any proton-donor, e.g. $\overset{+}{OH_3}$, $\overset{+}{NH_4}$, H_2O , HOEt, HO·CO·CH₃; but the negative pole, instead of being merely a proton-acceptor, has to supply hydroxyl ions to the ester, and so must actually contain the hydroxyl group. Nitrogenous bases can therefore only act indirectly by liberating hydroxyl ions from water molecules; and in the same way acetate anions can only provide hydroxyl by liberating it from water:

$$CH_3CO \cdot O + H_2O \Rightarrow CH_4CO \cdot O \cdot H + OH$$

The latter conclusion has been confirmed by the experiments of H. M. Dawson, who has shown that the catalytic coefficient of the acetate anion in the hydrolysis of ethyl acetate is only 1/30,000 of that of the oxonium ion.

The mechanism of hydrolysis which has been suggested above depends on bringing together (i) a molecule of the ester, (ii) a hydrogen or a hydroxyl ion of the acid or alkaline catalyst, and (iii) a molecule of water, since a hydrogen or a hydroxyl ion cannot readily be provided from any other source in acid and alkaline solutions, respectively. The operative stage in the hydrolysis, however, is not the mere formation of a complex (which might dissociate into the same substances from which it was formed), but a flow of electrons from atom to atom in the complex, whereby the carbon-oxygen bond of the ester is broken and an ion of the catalyst is released for further action.²

The rates of the three essential stages are presumably as follows:

⁽i) Addition of H or OH to the ester to form a complex ion. Slow, on account of the small concentration of the catalyst

¹ J. Chem. Soc., 1927, 2446. ² Lowry, J. Chem. Soc., 1927, 2558.

- (ii) Collision of the complex ion with a molecule of water. Rapid, in aqueous solutions, on account of the high concentration of water molecules.
- (iii) Rupture of the bond by a flow of electrons. Probably instantaneous, since it is analogous to an electrolysis in which the catalyst and the water act as poles between which the ester molecule is electrolysed.

These conclusions are supported by the calculations of Moelwyn-Hughes, who has shown that the velocity of hydrolysis of ethyl acetate and similar substances, in the presence of an acid, is approximately equal to that calculated on the supposition that chemical change occurs upon every ternary collision between an activated molecule of an ester, a hydrogen ion, and a water molecule.

The rate of hydrolysis is dominated by the concentration, or activity, of the catalyst, and by the stability of the complex ion, but may be influenced by the concentration or activity of the water when the hydrolysis is carried out otherwise than in dilute aqueous solutions.

The preceding mechanism is applicable, not only to hydrolysis, but also to prototropic change (see Chapter XIII), which again appears to depend on a flow of electrons through the prototropic molecule between a proton-donor and a proton-acceptor. Since the fundamental process in each case is the passage of an electric charge between two opposite "poles", provided by an amphoteric medium, it has been described by Lowry 2 as an electrolytic theory of catalysis by acids and bases.

(d) Esterification. The reverse process of esterification only takes place in presence of acid catalysts, since sodium acetate, for instance, cannot be used as an acetylating agent for alcohols. It may be formulated in a converse way to the process of hydrolysis, by postulating the formation of a ternary addition-compound of the acid catalyst with the organic acid and the alcohol, followed by the

¹ Phil. Mag., 1932, 14, 112; "The Kinetics of Reactions in Solution", pp. 249-52. See also Waters, J. Chem. Soc., 1936, 1014; Newling and Hinshelwood, J. Chem. Soc., 1936, 1357.

¹ J. Chem. Soc., 1927, 1554.

separation of a molecule of water and the liberation of the catalyst. The proton of the catalysing acid appears to act (i) by repressing the ionization of the organic acid, and consequently developing its ketonic functions 1; (ii) by forming an addition-complex with the alcohol,2 thereby imparting proton-donating properties to the otherwise inert molecules of the alcohol. The alcohol-complex can then unite with the carbonyl group of the acid, in accordance with the normal rules for the addition of hydrides, such as HCN to an aldehyde or ketone. Finally, the resulting complex must split off both the catalysing proton and a molecule of water in order to yield a molecule of the ester. The essential stages in the process may therefore be formulated in the following scheme, where the flow of electrons which accompanies the making and breaking of the bonds is indicated by broken arrows.

$$CH_{3} \xrightarrow{O} C_{2}H_{6} \rightarrow CH_{3} \xrightarrow{O} C_{3}H_{6} \rightarrow CH_{3} \xrightarrow{O} C_{4}H_{6} \rightarrow CH_{3} \xrightarrow{O} C_{4}H_{6} \rightarrow CH_{3} \xrightarrow{O} C_{5}H_{6} \rightarrow CH_{5} \rightarrow CH_{5} \xrightarrow{O} C_{5}H_{6} \rightarrow CH_{5} \rightarrow CH$$

Velocity of Esterification and of Hydrolysis.

Esterification and hydrolysis can both be attributed to the reversible formation of a ternary complex, e.g. of a molecule of the ester with a molecule of water and with an ion of the catalyst. The efficiency of any given catalyst is therefore determined largely by its ability to form a complex of this type. The fact that hydrolysis is arrested almost completely in neutral solutions shows that the intermediate complex is too unstable to persist in the form of a neutral bipolar molecule in the absence of acid or alkali.

Since the formation of this intermediate-product will

¹ Lowry, J. Chem. Soc., 1923, 123, 827; 1925, 127, 1381.

² Cf. Goldschmidt and Udby, Zeit. phys. Chem., 1907, 60, 728.

depend very largely on the nature of the medium, etc., it is necessary to standardize the condition of reaction very rigorously before any measurements of the velocity of hydrolysis or of esterification can be regarded as strictly comparable. This can be done (i) by using one ester and varying the concentration and nature of the hydrolysing medium or (ii) by comparing the velocities of hydrolysis of a series of esters in the same medium.

(a) Hydrolysis of Esters as a Basis for Measurements of Catalysis. The influence of environment upon the velocity of hydrolysis of an ester has been studied very extensively, since it provides one of the best methods of investigating the dynamics of homogeneous catalysis. The early measurements of Arrhenius and of Ostwald were interpreted on the hypothesis that catalysis in aqueous solution was due exclusively to hydrogen and hydroxyl ions. In 1889, however, Arrhenius observed that the rate of hydrolysis of aqueous solutions of sugars, esters, etc., by weak acids was augmented by the addition of neutral salts of strong acids, such as potassium chloride.1 He therefore suggested that the degree of ionization of weak acids may be increased by the addition of neutral salts. Subsequent workers have suggested (i) that the ions of neutral strong electrolytes decrease the dissociation of weak electrolytes present and so increase the concentration of catalysing molecules,2 or (ii) that the only real catalyst is a trace of free unsolvated hydrogen ions, derived either from solvated hydrogen ions or from undissociated molecules of any acid catalyst, as for example from molecules of acetic acid.3 This latter conception is in accord with the mechanism for ester-hydrolysis which has been put forward above, although it does not provide a complete explanation of hydrolysis in both acid and alkaline media.

¹ Zeit. phys. Chem., 1889, 4, 226.

² Goldschmidt, Z. Elektrochem., 1912, 18, 539.

⁸ Lapworth, J. Chem. Soc., 1908, 93, 2163, 2187; cf. Rice, J.A.C.S., 1923, 45, 2808.

(b) Catalysis by Hydrogen- and Hydroxyl-donors Arrhenius 1 also made the important observation that the effectiveness of an acid in catalysing the hydrolysis or "inversion" of cane-sugar was greatest in concentrated solutions. The catalytic activity of the acid therefore increased as its ionization decreased. Arrhenius was not willing to give up the hypothesis that this hydrolysis was catalysed exclusively by hydrogen ions, since this would have invalidated one of the principal methods then in use for determining their concentration. He therefore corrected his measurements by introducing a factor to represent the variable activity of the hydrogen ions at different concentrations, this factor increasing from 0.75 at 0.0005 N, to unity at 0.1 N and to 1.29 at 0.4 N concentrations.² The obvious interpretation, that the undissociated molecules of a strong acid can also act as catalysts, and are indeed even more effective catalysts than the ions, was suggested in 1912 by Snethlage,3 who concluded that the molecules of hydrochloric acid are twice as effective as its ions.

The view that the molecules of a strong acid as well as its ions are active catalysts was widely advocated as the "Dual Theory of Acid Catalysis", by H. M. Dawson, who subsequently extended it by showing that the anions of a weak acid can act as catalysts for reactions which are brought about by proton-acceptors. Subsequent observations have shown further that, whilst metallic kations have no marked catalytic activity, the kations of a weak nitrogenous base, e.g. NH₄+ have a marked catalytic activity in virtue of their behaviour as proton-donors.

(c) Kinetics of Catalysis by Acidic and Basic Molecules and Ions. H. M. Dawson has deduced the following equation for a reaction which is catalysed by hydrogen

¹ Zeit. phys. Chem., 1889, 4, 226.

² Ibid., p. 246.

³ Z. Elektrochem., 1912, **18**, 539; cf. Goldschmidt, loc. cit., and McBain, J. Chem. Soc., 1914, **105**, 1517.

⁴ J. Chem. Soc., 1915, 107, 1426.

⁵ Lowry and Smith, *J. Chem. Soc.*, 1927, 2539.

and hydroxyl ions, by undissociated molecules of acid, and by the anions of the acid.

Let v_H , v_{OH} = partial reaction-velocities of hydrogen and hydroxyl ions of concentrations [H] and [OH],

 v_M = partial reaction-velocity of molecules of concentration [HA],

 v_A = partial reaction-velocity of anions of concentration [A⁻]

then

 $v = v_H + v_M + v_A + v_{OH} = k_H[H^+] + k_M[HA] + k_A[A^-] + k_{OH}[OH^-]$ (1) where k_H , k_{OH} , k_M and k_A are the "catalytic coefficients" of the individual ions and molecules. For a weak acid catalyst at concentration c we can substitute:

$$[A^{-}] = [H^{+}], \quad [HA] = c - [H^{+}], \quad [H^{+}] \times [OH^{-}] = K_{w}$$

 $K_{a} = [H^{+}] \times [A^{-}] \div [HA]$

where K_{w} , K_{a} are the dissociation constants of water and of the weak acid. We can thus obtain an equation, containing only two variables, that is capable of direct mathematical solution.

In acid solutions (of $p_{\rm H}=3$ or under) the catalytic effect of the minute trace of OH⁻ ions is negligible. Equation (1) may therefore be written in the form: ²

$$v = (k_H - k_M)[H^+] + k_a \frac{K_a c}{[H^+]} + k_M c - k_a K_a$$
 . (2)

This is a quadratic equation and gives a minimum velocity v_i at a concentration $[H_i^+]$ of hydrogen ions as represented by the following formulæ:

 $[H_i^+] = \sqrt{k_a K_a \cdot c \cdot / (k_H - k_M)}$ and $v_i = 2 \sqrt{(k_H - k_M) k_a K_a c} + k_M c$ In this mixture of minimum velocity the induced catalytic effects of the positive and negative ions $(v_{Hi}$ and v_{Ai}) are equal, and their value is given by the expression

$$\sqrt{(k_H - k_M)k_a \cdot K_a \cdot c} = \frac{1}{2}u_i$$

If the reaction-velocity of any system which is catalysed by both positive and negative ions is examined, this type of

¹ J. Chem. Soc., 1926, 2282. ² Ibid., 1926, 2872, 3167.

relationship appears to hold true. In the case of a reaction like ester hydrolysis, which is catalysed by both hydrogen and hydroxyl ions, the reaction-velocity, if plotted against

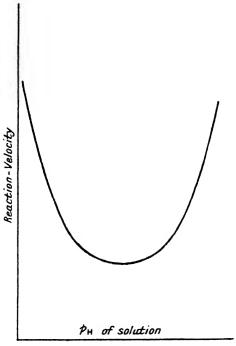


FIGURE VI.

the hydrogen ion concentration (or $p_{\rm H}$) of the solution gives a symmetrical curve having a minimum value.

The reaction-velocity curve of the type of Figure VI is a member of a sheaf of catenaries, which are described by Dawson as catalytic catenaries, the parameters of the catenary being determined by the nature and concentration of the added catalysts. In particular it can be seen from the equations already given that the concentration of hydrogen ions $\begin{bmatrix} + \\ H_i \end{bmatrix}$ at which the reaction-velocity falls to a minimum is proportional to the square root of the concentration of catalysing acid.

A simpler equation than (2) for the representation of these catenaries may be derived by writing

$$u = v_h + v_a$$
 and $[H^+] = n[H_i^+]$.

Then it is seen that

$$\frac{u}{u_1} = \frac{1}{2} \left(n + \frac{1}{n} \right) = r; \text{ or } r = \frac{1}{2} (10^{\Delta p_{\text{M}}} + 10^{-\Delta p_{\text{N}}})$$

which is of the form $r = \cosh h$, where $h = \log_e n$.

This equation, the reduced catalytic catenary, applies equally well to catalysis by water as to catalysis by acids or alkalis, and has been applied with great success to the hydrolysis of esters.2 However, for ester hydrolysis in water, or even in buffer solutions containing dissolved salts, account must be taken both of the catalytic effect (constant) of any other dissolved ions that may be present. and of the autocatalytic effect of the hydrogen ions and organic acid anions formed from the ester itself as the reaction proceeds. It may be mentioned that, in accordance with anticipations from the modern theory of strong electrolytes, inorganic salts at concentrations of o r N or less can be assumed to be completely dissociated and consequently to act only as separate ions and not as "ion pairs" or salt molecules. When these factors are fully taken into account the discrepancies and the puzzling variations of reaction-velocity noted by the earlier workers on hydrolysis can be explained, and a constant correlation coefficient between reaction-velocity and ionic (or molecular) catalyst concentration (not "activity") can be calculated.

Influence of Structure on Velocity of Hydrolysis of Esters.

(a) Influence of Polar Groups. By making use of Dawson's equations, Ingold 3 has been able to correlate the

¹ J. Chem. Soc., 1927, 213.

² Dawson and Lowson, J. Chem. Soc., 1927, 2107, 2444; 1928, 2146, 3218; 1929, 393, 1229; cf. Wijs, Zeit. phys. Chem., 1893, 11, 492.

⁸ J. Chem. Soc., 1930, 1032.

catalytic coefficients of hydrogen and hydroxyl ions with the polar nature of the groups, G and R, in the hydrolysis of an ester, G·CO·O·R. He considers that in any "sotypic" series of esters, G·CO·O·R, in which geometrical or "steric" variations of the groups G or R cannot affect reaction velocity, the ratio $C = k_{\rm OH}/k_{\rm H}$ should be a function of the polarity of the groups G and R. The variations of this ratio, in two series of esters, is shown in the following table.

TABLE I

Hydrolysis of Esters by Acids and Alkalis

(a) Esters of series X·CH₂·CO·O·CH₂·CH₂·Y

X	Y	ри for Minimum Velocity.	Temp.	K _w 10 ¹⁴	$k_{\mathrm{OH}}/k_{\mathrm{H}} \times 10^{-8}$
H H Cl CH ₃ CO NH ₃ +	Н НО Н Н	5·36 5·0 4·8 4·4 3·5	25° 25° 20° 25° 10°	1.0 1.0 0.8 1.0 0.8	Effect of "electronsink" groups

(b) Esters of series X·CO·O·CH₂·Y

H	H				large	
H	CH ₃	4.65	20°	0.8	63)	Effect of
CH_3	Н	5.15	25°	1.0	5	" electron-
CH ₃	CH ₃	5.36	25°	1.0	1.9	source "
CH_3	CH ₃ ·CH ₂ ·CH ₂	5.65	25°	1.0	0.5)	groups
	1			1	Į.	

It can be seen that "electron-sinks", i.e. groups producing positive fields, C-R, favour catalysis by hydroxyl ions, whilst electron-sources, producing negative fields, C-R, favour catalysis by hydrogen ions. In other words, acylous groups like Cl, which produce a negative general polarity (-I) or a positive field (+F), favour the formation of the anion A, by attracting hydroxyl ions. On the

¹ This would be satisfied, for example, in any group of esters such as

other hand, basylous groups like CH_3 , which produce a positive general polarity (+I) or a negative field (-F) favour the formation of the kation B, by attracting hydrogen ions from the medium. This relationship holds good, whether the polar group forms part of the acidic or of the alcoholic radical of the ester.

It would thus appear that the mechanism of ester hydrolysis outlined above (p. 268) can be extended by saying that the polarity of the radicals (G, R) constituting the ester G·CO·O·R is responsible for attracting catalysing ions (H⁺ and OH⁻) from the aqueous solution, so as to produce the activated bipolar molecule (C) which eventually breaks down into the constituent acid and alcohol, and at the same time liberates once more the ion of the catalyst.

(b) Relative Reactivities of Alkyl Glycerates. Direct measurements of the relative catalytic coefficients of hydroxyl and of hydrogen ions $(C = k_{\rm OH}/k_{\rm H})$ in the hydrolysis of esters have been made for simple alkyl groups forming the alcoholic fraction (R) of esters of glyceric acid $HQ\cdot CH_1\cdot CHOH\cdot CO\cdot O\cdot R$.

TABLE II

* Hydrolys	is of Alk	γl	Glycerates	by	Acids a	nd Alkalis 1
Ester.				þн	(min. rate).	$k_{\rm OH}/k_{\rm H} \times 10^{-3}$.
Methyl dl-	glycerate				4.165	27.5
Ethyl	,,		•		4.320	13.5
n-Propyl	,,		•		4.350	11.8
iso-Propyl	,,		•		4.505	5⋅8
n-Butyl	,,		•		4.355.	11.5
iso-Butyl	,,				4.370	10.7
n-Amyl	,,				4.360	11.2
iso-Amyl	**	•	•	•	4.360	11.2

¹ Groocock, Ingold and Jackson, J. Chem. Soc., 1930, 1057.

These results indicate that alkyl groups favour catalysis by hydrogen ions, but that this effect shows no further increase when the growing chain contains 3 or more atoms of carbon. This limitation is in harmony with the results already recorded for other manifestations of general polarity (cf. Chapter XI).

(c) Hydrolysis of Esters of Dibasic Acids. The magnitude of the influence of general polarization on the velocity of hydrolysis of esters of dibasic acids has been discussed by Ingold 1 on the lines originated by Bjerrum in dealing with the dissociation constants of dibasic carboxylic acids (p. 214). For esters of symmetrical dibasic acids the rates of hydrolysis of the first ester grouping (k_1) should, on statistical grounds, be twice that of the second ester group (k_2) . This was observed by J. Meyer in 1909 2 and discussed more fully by Wegscheider in 1915.3 It holds true. however, only for acid hydrolysis of esters, and is not even approximately true for alkaline hydrolysis or saponification.4 An explanation may be found in the view that polar influences originating within the ester molecule will assist or hinder reaction by attracting or repelling the catalysing H⁺ and OH⁻ ions of the reaction mixture and can thus alter the relative speeds of hydrolysis of the two ester groups of dicarboxylic acids. This is of dominating importance in alkaline hydrolysis, which is catalysed chiefly by hydroxyl ions, since the residue, R·O·CO·(CH₂)_n·COO⁻, which remains when the first ester grouping has been saponified, is an anion and will necessarily repel the hydroxyl ions from the vicinity of the second ester grouping, and thus depress the rate of hydrolysis k_* still further below the value $\frac{1}{2}k_1$.

Ingold calculated the effect of the $-COO^-$ pole upon the rate of saponification of the second ester group of a half-hydrolysed dicarboxylic ester, of the type

¹ J. Chem. Soc., 1930, 1375.

² Zeit. phys. Chem., 1909, 66, 81.

³ Monatshefte, 1915, 36, 471.

⁴ Zeit. phys. Chem., 1909, 67, 257.

 $RO\cdot CO\cdot (CH_1)_n\cdot CO\cdot OR$ (in the same way as Bjerrum calculated the effect of a $-COO^-$ pole upon the dissociation constant of the second hydrogen ion of a dibasic acid), by evaluating the work required to bring up the ionic catalyst against the external electrical field of the $-COO^-$ pole. For a symmetrical dicarboxylic ester undergoing alkaline hydrolysis this calculation leads to the relationship:

$$k_1/k_0 = 286.97 \times 10^{-8}/_{p}$$

where k_1 and k_2 are the first and second hydrolysis-constants for alkaline solutions. The intramolecular distances, r, calculated by this equation agree with those deduced by Gane and Ingold ¹ from measurements of the dissociation constants of dibasic acids (p. 216).

TABLE III

Hydrolysis of Esters of Dibasic Acids

						7 × 10	
Ester Hyd:					k_1/k_2 .	From Ester.	From Acid. 1
Oxalate	methyl	•	•		104	r	
	ethyl				104		
Malonate	methyl				92		
	ethyl				91	r·8	1.5
Succinate	methyl				9.6		
	ethyl	•	•		9.9	4.4	5.0
Glutarate	methyl	•	•		6⋅8		
	ethyl	•			6∙1	6∙o	9.2
Suberate	methyl		•		3.1	15	14.5
Azelate	methyl	•	•		2.95	17.5	17
Sebacate	methyl	•	•	•	2.8	20	

When catalysis by both hydrogen and hydroxyl ions is taken into consideration, the equation becomes:

$$\frac{(k_1)_{\text{OH}}/(k_2)_{\text{OH}}}{(k_1)_{\text{H}}/(k_2)_{\text{OH}}} = e^{6.97 \times 10^{-8}/r}$$

This equation is quite consistent with that given above, because the ratio $(k_1)_{\rm H}/(k_2)_{\rm H}$ has the constant value 2.0 ³ for the esters of oxalic acid and its higher homologues. When applied to the catalytic catenary of a dicarboxylic ester, it leads to an approximate relationship between the

¹ J. Chem. Soc., 1928, 1594. ² Ibid., 1930, 1380.

³ Palomaa, Ann. Acad. Scient. Fennicæ, 1917, A.10, No. 16; Skrabal, Monatshefte, 1917, 38, 29.

intramolecular distance r and the hydrogen-ion concentrations, p_{H}' and p_{H}'' , of minimum reaction velocity for the first and second stages of hydrolysis, as follows:

$$p_{\rm H'} - p_{\rm H''} = \pm 3/10^8 r$$

The sign of the numerator is chosen to correspond with that developed on the ionogenic ester.

Ingold's equations can be applied to the hydrolysis of esters by water or by very dilute electrolytes, but not to hydrolysis in concentrated solutions of "strong electrolytes", where the electrostatic fields due to other ions (e.g. the Na⁺ in strong NaOH) will affect the local gradients of electrical potential in the vicinity of the molecules of ester, and will consequently influence the work involved in bringing a catalysing ion into contact with the ester group. Also, as in Gane and Ingold's measurements of dissociation constants of acids, internally-propagated polarity should be taken into consideration as well as the directly-transmitted effect, and no data are yet available for the computation of these influences.

- (d) Interpretation of Earlier Measurements. The preceding ad hoc study of the influence of a COO⁻ pole upon the velocity of hydrolysis of an ester group illustrates a modern quantitative method of investigation, which has not yet been applied very widely. A great many measurements of rates of esterification of acids, and of hydrolysis of esters, have been recorded, however, during the past fifty years, and these provide valuable material for a qualitative study of the influence of various polar groups on this type of chemical change.
- (i) Reicher i in 1885 showed that the alteration of the acyl residue in an ester affected its rate of saponification far more than did alteration of the alcoholic residue. Thus methyl acetate was saponified about 50% more rapidly than ethyl acetate; but the further decrease in the velocity of saponification of the esters of the higher alcohols was very slight. The acetate esters of secondary and

tertiary alcohols, however, were more difficult to saponify than those of primary alcohols.

(ii) In 1894 Victor Meyer drew attention to the great practical difficulty of preparing certain ortho-substituted aromatic esters. Thus, mesitylene carboxylic acid (I) could not be esterified by the Fischer-Speier method, with hydrogen chloride as the catalyst, whereas mesitylenic acid (II) and mesityl acetic acid (III) both yielded esters quite readily. Nevertheless, the silver salt of any organic acid would give the ester on treatment with ethyl iodide:

COOH

CH₃

CH₃

CH₃

$$CH_3$$
 CH_3
 C

A more detailed investigation by Victor Meyer and Kellas ² of the relative rates of esterification of substituted aromatic acids by the Fischer-Speier method, by a standardized experimental procedure, showed that (in the presence of the acid catalyst) nearly all substituted aromatic acids were esterified less readily than benzoic acid itself, but that the *ortho*-substituted acids were definitely the least reactive.³ They therefore formulated the following *esterification law*:

"When the hydrogen atoms in the two ortho positions to the carboxyl in a substituted benzoic acid are replaced by radicals such as Cl, Br, NO₂, CH₃, COOH, an acid results which can only be esterified with difficulty or not at all."

The following figures, due to Kellas, giving the percentage of acid esterified in two hours at 51° illustrate the results obtained.

¹ Ber., 1894, 27, 510, 1580.

² Zeit. phys. Chem., 1897, 24, 221. ³ Ibid., 1897, 24, 219.

284 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

TABLE IV

Esterification of Substituted Benzoic Acids

Substitu	ent.				- CH _a	Cı	Br	1	NO _s	
ortho			•	•	48·3 77·1 75·6	50.9	43.4	20.5	8.6	Benzoic acid
meta	•	٠	•	٠	77·I	72.0	66.6	57.6	57·I	82.5
para	•	•	•	•	75.0	70.5	01.0	52.9	57·I	3

(iii) The esterification law, which forms one of the main pillars of the theory of steric inhibition of chemical change, was extended to the aliphatic series by Sudborough, who showed that the cis-isomers of the unsaturated acids of the substituted acrylic acid series CHX=CY·COOH, were more difficult to esterify than their trans-isomers,1 and that the introduction of substituent alkyl groups greatly reduced the speed of esterification of succinic acids.2 More detailed investigations by Sudborough, Reid, and others, however, have indicated that steric hindrance by ortho-substituents is only one of the many factors affecting velocity of esterification or of hydrolysis. Thus changes in conditions of reaction, which are equivalent to changing the activating catalyst, can alter the entire order of reactivity throughout a series of substituted acids, e.g., esterification by the Fischer-Speier process, in presence of hydrochloride as a catalyst, occurs less readily with a substituted aromatic acid than with benzoic acid; but esterification in the absence of a catalyst proceeds most readily with the substituted acids.³ Similarly, the saponification of an ester is accelerated by a halogen, nitro or carboxyl group even in the ortho-position, but is retarded by methyl groups.4 It is, therefore, clear that although "steric hindrance" must be taken into consideration when dealing with vicinal substituents, the rate of reaction is usually dominated by

¹ Sudborough and Lloyd, J. Chem. Soc., 1898, 73, 81.

² Bone, Sudborough and Sprankling, J. Chem. Soc., 1904, 87, 534.

³ Michael, Ber., 1909, 42, 317.

⁴ Blakey, McCombie and Scarborough, J. Chem. Soc., 1926, 2863.

influences arising from the polarity of the substituent groups.¹

The different orders of induced reactivity recorded under different sets of experimental conditions are principally due to changes in the acidity or alkalinity of the medium. Thus, the retardation of esterification of ortho-substituted acids was discovered by experiments in which an acid catalyst was used, and does not occur when esterification is effected in neutral solutions by the action of ethyl iodide on the silver salt of the acid. On the other hand, the acceleration of saponification by certain substituents was necessarily observed in alkaline solutions. This reversal of sign is easily interpreted, however, by noticing that groups, such as the halogens, which are electron-attracting, would inhibit catalysis by hydrogen ions, but would promote catalysis by hydroxyl ions, whilst the converse would hold for groups which repel electrons.

The relevant data, which are set out in Tables V to VII, are therefore in good agreement with the theory of polarity dealt with in the preceding chapter. Table VII is of particular interest because it shows that in the saponification of esters, as in other reactions, polar effects are cumulative, and that the rate of reaction of a di-substitution-derivative can be determined from that of a monosubstitution derivative by multiplicative equations of the type:

 $k_{2^{-4}-ester}/k_{benzoate} = (k_{2-ester}/k_{benz.}) \times (k_{4-ester}/k_{benz.})$

This latter relationship indicates that the effects of substituent groups on the energies of activation of ester groups are additive.

Moreover, Nathan and Watson³ have shown that the activation energies for alkaline saponification of substituted aromatic esters, which is a *Class B* reaction (p. 243), satisfy the equation

$$E = E_0 + a\mu + b\mu^2$$

¹ Hinshelwood and Legard, J. Chem. Soc., 1935, 595, 1588.

² Cf. Olivier, Rec. trav. Chim., 1926, 45, 296, Ostwald, Zeit. phys. Chem., 1889, 3,, 385, Hollemann, Rec. trav. Chim., 1901, 20, 361.

³ Nathan and Watson, J. Chem. Soc., 1933, 1249.

286 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

by which they can be correlated with the magnitudes of the dipole moments of the substituent groups (see p. 248).

Table V Velocity of Saponification of Ethyl Esters in Aqueous Alcohol at 30° C. Relative velocities calculated as K(ester)/K(bensoate)

				С	innamates	(k).			
Substituent.			nt.	ortho.	meta.	para.	ortho.	meta.	para.
CH ₃		•	(m) (k)	0.584	0·746 0·704	0·506 0·467			1.54
F.			(<i>k</i>)	5.55		1.916	5.06		2.96
Cl.		•	$m \choose m' \choose k$	2·89 2·54 1·91	8·93 6·68 7·41	3·83 4·14 4·33			4:39
Br			(m) (k)	2·17 1·89	8·20 8·07	4·31 4·93	6.63		4.85
Ι.	•	•	$\binom{m}{k}$	1·63 0·835	7.51	4·74 5·08	5.45		4.88
NO ₂			(k)	57.1	62-11	104.2		16.3	21.2
CH ₃ O			(m')	1.30	1.25	0.225			
CH ₃ S			(k)			0.214			1
NH ₂	•	•	(k)		0.422	0.023	(Cinn	1·35 amic ac 2·16)	cid =

⁽m) = calc. from McCombie and Scarborough, J. Chem. Soc., 1915, 107, 156 (98% EtOH).

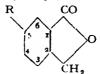
The results for the cinnamates are of special interest because they illustrate the elimination of the effects of steric hindrance of the ortho-substituents in the aromatic ring (cf. Kindler, Annalen, 1928, 464, 278)

⁽m') = calc. fromBlakey, McCombie and Scarborough, J. Chem. Soc., 1926, 2863 (95% EtOH).

⁽k) = calc. from Kindler, Annalen, 1926, 450, 3; 1927, 452, 90; 1928, 464, 278 (87.83% EtOH).

TABLE VI

Velocity of Hydrolysis of 5-substituted Phthalides 1



Phthalide deriv.				K.:	ro4 at 25°.	K.subs./K.unsubs.
Phthalide .	•	•	•		49	1.000
Chlorophthalide					65	1.33
Bromophthalide	•		•		75	1.53
Iodophthalide			•		98	2.00
Cyanophthalide		•	•		280	5.7
Nitrophthalide					520	10.8
Methoxyphthalide	(for	meco	nine)		32	0.65
Aminophthalide '					15	0.31
Acetylaminophtha	lide	•			37	0.75

TABLE VII

Saponification of Substituted Ethyl Benzoates 2

(a) E:	ster.			k in 70% Alcohol.	k in 95% Alcohol.
Ethyl b	enzoate .	•		0.105	0.028
Ethyl-o	-methoxy-b	enzoate		0.077	0.0365
,, -11	1-methoxy-	**		0.126	0.035
,, -p	-methoxy-	**		0.023	0.0063
,, -0	-chloro-	**		{0·116 0·116	0·096 0·071
,, -11	ı-chloro-	,,	•	0.74	o·187
., -p	-chloro-	,,		0.40	0.119

(b) Ethyl Dimethoxy-benzoates 3

Ester.	kobs.	keste. 95% Alcoho	kobs./konic.	kobs.	keale.	kobs./koale.
125101.			••			
2:3.	0.001	0.046	1.33	0.110	0.093	1.18
2:4.	0.0084	0.0082	1.02	0.010	0.017	1.12
2:5.	0.057	0.046	1.24	0.122	0.093	1.31
3:4 .	0.0102	0.0079	1.29	ი∙ივ8	0.028	1.35
3:5.	0.046	0.044	1.05	0.175	0.121	1.16
-	(c) 1	Ethyl Di	chloro-ben	zoates		
2:3.	0.157	0.474	0.33	0.34	0.817	0.42
2:4 .	0.206	0.295	0.71	o·65	0.44	1.47
2:5.	0.42	0.474	0∙89	1.04	0.817	1.27
3:4 .	0.57	0.775	0.72	1.95	2.82	0.70
3:5 •	1.52	1.25	1.22	4.20	5.21	0.81

¹ Tasman, Rec. trav. Chim., 1927, 46, 653-98.

3 Ibid., 1926, 2865.

² Blakey, McCombie and Scarborough, J. Chem. Soc., 1926, 2867

CHAPTER XIII

IONOTROPIC CHANGE

Isomerism and Isomeric Change

The term isomerism (Greek, loos, equal; $\mu \ell \rho o s$, a part) was introduced by Berzelius in 1831 ¹ in order to describe the phenomenon of identity of composition in substances which were obviously different from one another in their properties. This phenomenon was first disclosed in 1824 when Wöhler's analysis of silver cyanate ² and Liebig and Gay-Lussac's analysis of silver fulminate ³ gave results which led to identical formulæ for the two salts, ⁴ both of which contained 72·1% Ag; but Berzelius was not convinced of the reality of these observations until his own experiments ⁵ had confirmed the statement, made by Gay-Lussac in 1826, ⁶ that racemic acid (discovered by Kestner in 1822) was identical in composition with the tartaric acid of Scheele.

Greater interest attaches to Wöhler's discovery that ammonium cyanate could be converted into urea, merely by evaporating to dryness a solution prepared by the action of ammonia on lead cyanate. This discovery broke down the barrier between mineral and vital chemistry, by showing that a typical product of animal metabolism could

¹ Pogg. Ann., 1830, 19, 326; Jahresber., 1831, 11, 44.

² Ann. Chim. Phys., 1824, [2], 27, 196.

⁸ Ibid., 1824, [2], **25**, 285; ibid., **24**, 294.

⁴ Ibid., 1824, [2], 27, footnote, pp. 199-200.

⁵ Jahresber., 1826, 6, 102.

⁶ Paper read at the Académie des Sciences, on Nov. 22nd, 1826, and abstracted in Schweigger's Journal, 1826, 48, 381; cf. also, Walchner, ibid., 1827, 49, 238.

⁷ Pogg. Ann., 1828, 12, 253.

be prepared in the laboratory from inorganic materials, namely ammonia, potassium cyanide and lead oxide. It is, however, also of interest as the first recorded example of *isomeric change*, i.e. of the conversion of one of a pair of isomeric compounds into the other.

Under the conditions used by Wöhler, a complete conversion of ammonium cyanate into urea was effected. Seventy years later, however, Walker and Hambly 1 showed that the change, which passes to completion when the solution is evaporated to dryness, is reversible in the liquid phase, and therefore affords an example of reversible isomeric change. Thus in a normal solution at 59.6° , 14.4% of ammonium cyanate and 0.0038% of urea undergo isomeric change in a minute. Equilibrium is reached when there is 5% of ammonium cyanate to 95% of urea:

$$[NH_4][CNO] \xrightarrow{95} CO(NH_2)_2$$
5%
95%

since the velocity of change of the ammonium salt is reduced to 7.2% per minute at its final concentration of about N/20, and 7.2/0.0038 = 95/5. The reduction in the velocity of change of ammonium cyanate in dilute solutions tends to stabilize the salt, so that at a concentration of N/2,000 the proportions of ammonium cyanate and urea in the equilibrium mixture would be approximately equal.

This isomeric change obviously depends on the fact that isocyanic acid can combine with ammonia in two ways:

(i) To form a salt
$$NH_1 + HNCO = NH_4N:C:O$$

OH

(ii) To form an amide $NH_1 + HNCO = HN:C$
 NH_2

Compare $NH_3 + CH_3:CHO = CH_3:CH$
 OH
 OH

If these two actions are reversible, so that ammonia can

1 Trans. Chem. Soc., 1895, 67, 746.

dissociate both from the salt and from the amide, a condition of equilibrium will at once be established between the two isomerides. It is of interest also to note that this method of formulation gives rise to Werner's formula for urea 1 instead of the traditional "carbamide" formula.

Dynamic Isomerism. Only slow progress was made in the study of isomerism and isomeric change until the development of the doctrine of valency made it possible to assign definite structures to the molecules of organic compounds. The facts in reference to the isomerism of racemic and tartaric acids, which had been established by the experiments of Pasteur, were then elucidated almost immediately on the basis of van't Hoff and le Bel's theory of the tetrahedral carbon atom; and the theory of reversible isomeric change was established on a firm experimental basis by the publication in 1877 of Butlerow's work on "iso-Dibutylene".2

(a) Interconversion of Butylene and Butyl Alcohol. Butlerow's discovery was made in the course of a series of experiments on the conversion of the hydrocarbon butylene, C₄H₈, by the action of sulphuric acid, on the one hand into butyl alcohols of the formula C₄H₁₀O, and on the other hand into polymeric olefines such as C₈H₁₆ and C₁₂H₃₄. This investigation may be regarded as an extension of the work of Faraday on the relationship between ethylene and butylene, and formed part of a general study of the polymerization of the olefines and their interconversion with the related alcohols.

As early as 1873 Butlerow had noticed the presence of CH₃.

isobutyric acid, CH.CO:OH, amongst the products

obtained when carefully purified crystalline specimens of trimethyl carbinol, C(CH₂)₃OH, were oxidized by means of a mixture of bichromate and sulphuric acid. The formation of this acid was a clear indication that the liquid con-

¹ J. Chem. Soc., 1912, **101**, 2185; 1917, **111**, 863; 1919, **115**, 1093.

² Annalen, 1877, 189, 44.

tained, in addition to the original tertiary compound, a CH₃ CH·CH₂·OH.

That the two alcohols were closely related was already evident from observations made six years previously. The tertiary compound had been detected—as an impurity or as a product of change—in commercial isobutyl alcohol after conversion into the chloride, and an actual change from one alcohol into the other had been effected by converting the tertiary alcohol by the action of sulphuric acid (I vol. to 2 vols. of water) into isobutylene, combining this with hypochlorous acid and reducing out the chlorine with sodium amalgam—

A more direct conversion of isobutylene into a butyl alcohol, by absorbing it in sulphuric acid (2 vols. to I vol. water), diluting and distilling, had led to the reproduction of the original tertiary compound as the main product of the action, but no method was available for the ready separation of the non-crystalline primary isomeride, and no attempt was therefore made at the time to detect it as a minor constituent of the resulting alcohol, although the possibility of its formation was clearly recognized. When, however, it was found in 1873 that its presence in the products of interaction of tertiary butyl alcohol and sulphuric acid could be demonstrated by its ready oxidation to isobutyric acid, it was at once evident

¹ Annalen, 1867, 144, 34.

² Loc. cit., p. 30.

⁸ Loc. cit., p. 22.

that its formation must be ascribed to a reversible dehydration of the alcohol followed by recombination with water (or with sulphuric acid) in both of the two possible ways:

(b) Amylene and Amyl Alcohol. Although this important observation was made in 1873, it was not until four years later that Butlerow was able to make a clear pronouncement on the subject of reversible isomeric change. The paper, "Ueber Isodibutylen", in which his views were first fully developed, will always hold rank as a classic, and can be regarded as second in value only to the papers in which Liebig and Wöhler announced their discoveries of isomerism and of isomeric change half a century earlier.

A part of the paper is devoted to observations on the hydrocarbon amylene, C_8H_{10} , a compound intermediate in complexity between isobutylene and isodibutylene. Although very ready to undergo polymerization, the hydrocarbon could be prepared in a fairly pure form by heating tertiary amyl alcohol in a sealed tube at 100° with two or three volumes of a mixture of equal weights of sulphuric acid and water; under these conditions it separated as an oil on the surface of the acid, but when the mixture was allowed to cool and shaken repeatedly, the oil redissolved in the acid (by conversion into the alcohol or its sulphate), leaving only a small residue of diamylene, $C_{10}H_{20}$, on the surface:

$$\begin{array}{c|c}
HO \cdot C(CH_3)_3 & \longrightarrow & C(CH_3)_2 \\
CH_2 \cdot CH_3 & & H_2O \\
\text{Tertiary amyl} & & CH \cdot CH_3 \\
\text{alcohol.} & & Amylene.
\end{array}$$

By heating and cooling alternately these changes could be repeated to any desired extent. In this way the reversibility of the hydration and dehydration, postulated in the

case of isobutylene, was experimentally demonstrated as a general action which might be assumed to take place to a greater or less extent whenever an alcohol or an olefine was brought into contact with a sulphuric acid of moderate strength.

(c) iso-Dibutylene. iso-Dibutylene, the main subject of the memoir, was prepared by the action of 50% sulphuric acid on tertiary butyl alcohol. Its main oxidation products were acetone and trimethylacetic acid and its constitution was therefore:

This structure could be accounted for on the view that the actual condensation took place between a molecule of the tertiary alcohol and a molecule of *iso* butylene derived from it by the removal of water.

$$\begin{array}{c} CH_3 \\ CH_3 \\ \end{array} \longrightarrow \begin{array}{c} C=C-H+HO-C \\ CH_3 \\ CH_3 \\ \end{array} \longrightarrow \begin{array}{c} CH_3 \\ CH_3 \\ \end{array} \longrightarrow \begin{array}{c} CH_5 \\ CH_3 \\ \end{array} \longrightarrow \begin{array}{c} CH_5 \\ CH_3 \\ \end{array}$$

This structure was confirmed by combining the hydrocarbon with hydriodic acid, and hydrolysing the iodide by means of moist silver oxide. The product, "isodibutol". had the composition indicated by the formula C₈H₁₈O, and showed all the properties of the tertiary alcohol HO·C(CH₃)₂·CH₂·C(CH₃)₃, including a marked stability towards oxidizing agents. When the oxidation was forced by using a concentrated acid mixture, the products of the action were the unchanged alcohol, isodibutylene—formed by the dehydrating action of the sulphuric acid-and oxidation products derived from it: no direct oxidation of the alcohol could be detected. The experiments served, however, to show that the interconversion of olefine and tertiary alcohol, noticed in the case of butylene and amylene, took place also in the case of isodibutylene, and a more detailed study of the oxidation revealed the presence of an even more complex equilibrium involving not only two isomeric alcohols, but two isomeric olefines.

An oily acid left behind after the distillation of the trimethyl-acetic acid proved to be an octylic acid of the formula $C_8H_{18}O_2$, and the conditions of its formation left no doubt that it was the analogue of the *iso*butyric acid obtained as a by-product in the oxidation of tertiary butyl alcohol, and that its formation was due to a similar series of changes as set forth below:

$$(CH_3)_3C \cdot CH_4$$

$$CH_3$$

$$CH_3$$

$$CH_3$$

$$CH_3$$

$$CH_3$$

$$CH_3$$

$$CH_3$$

$$CH_3$$

$$CH_3$$

$$CH_4$$

$$CH_3$$

$$CH_3$$

$$CH_3$$

$$CH_4$$

$$CH_3$$

$$CH_4$$

$$CH_3$$

$$CH_4$$

$$CH_3$$

$$CH_4$$

$$CH_3$$

$$CH_4$$

$$CH_3$$

$$CH_4$$

$$CH_4$$

$$CH_4$$

$$CH_3$$

$$CH_4$$

$$CH_4$$

$$CH_4$$

$$CH_5$$

$$CH_5$$

$$CH_6$$

$$CH_7$$

$$CH$$

The remaining oxidation product was a ketone, C₇H₁₄O, which was oxidized by a stronger chromic acid mixture to acetic acid and trimethylacetic acid, and was therefore formulated as:

$$CH_3 \cdot CO \cdot CH_2 \cdot C(CH_3)_3$$
.

The ketone was evidently produced from the olefine

by oxidation of the =CH₂ group, and indicated the presence of this new hydrocarbon in the oxidizing mixture.

The complete equilibrium established by means of sulphuric acid acting on isodibutol or on isodibutylene may be represented by the scheme:

At ordinary temperatures the proportion of hydrocarbon would be very small, and the main action would be an equilibrium between two alcohols through the intermediate olefine,

$$\begin{array}{c} C(CH_3)_2 \cdot OH \\ | \\ CH_2 - C(CH_3)_3 \end{array} \longleftrightarrow \begin{bmatrix} C(CH_3) = CH_2 \\ | \\ CH_2 \cdot C(CH_3)_3 \end{bmatrix} \longleftrightarrow \begin{array}{c} CH(CH_3) \cdot CH_2 \cdot OH \\ | \\ CH_2 \cdot C(CH_3)_2 \end{array}$$

At high temperatures, on the other hand, the two olefines would predominate, but these would be brought into equilibrium by the presence of a small amount of the intermediate alcohol,

$$\begin{array}{c} \text{C(CH_3)_2} \\ || \\ \text{CH-C(CH_3)_3} \end{array} \xrightarrow{\longleftarrow} \begin{bmatrix} \text{C(CH_3)_2OH} \\ || \\ \text{CH_2-C(CH_3)_3} \end{bmatrix} \xrightarrow{\longleftarrow} \begin{array}{c} \text{C(CH_3)=CH_2} \\ \text{CH_2-C(CH_3)_3} \end{array} \\ \text{Isomeric} \\ \text{olefine.} \end{array}$$

The complete equilibrium between the four compounds could therefore be reduced, by varying the conditions, to either of two simple cases of reversible isomeric change.

(d) Isomeric Change without a Catalyst. It was one of the striking merits of Butlerow's observations that they not only established the existence of reversible isomeric changes, but also revealed the mechanism by which these changes were brought about. In the absence of the acid, or on diluting it with water, the hydration of the olefine and the dehydration of the alcohol were alike suspended. Under these conditions, therefore, dynamic equilibrium between isomers was no longer possible, and the isomerides became as stable as, for instance, ethyl alcohol and methyl ether. As Butlerow expressed it, a special reagent—sulphuric acid—was necessary to bring about in the tertiary alcohols the

"condition of equilibrium, depending on incessant isomeric change", which he had studied so successfully in the derivatives of isodibutylene.

But is such a special reagent always necessary in order to establish a reversible isomeric change? In Butlerow's opinion this was not so, and his discussion of this alternative type of change forms one of the most important features of his paper. Recent work has rendered it doubtful whether such cases actually exist; but a large number of instances are known in which a condition of equilibrium is established in presence of traces of alkaline or saline impurity so minute that it is a matter of the utmost difficulty to get rid of them, and unless extraordinary precautions are taken the state of affairs postulated by Butlerow is substantially realized —at least, to the extent that isomeric change proceeds without the conscious addition of a catalytic agent. The conclusions which he drew in reference to compounds which behaved in this way may be described most clearly by a direct paraphrase:

"In such cases, all attempts to establish a definite chemical structure for the substance would be uscless, since molecules of the two or more isomeric varieties would always be present. It is clear that the chemical metamorphoses of such a substance must proceed sometimes in the sense of one chemical grouping, and sometimes in the sense of the other, according to the nature of the reagent and the conditions of the experiment. As an example of such a two-sided chemical structure, one might probably adduce the cases of cyanic acid, hydrocyanic acid, and so forth. From this point of view it appears both unnecessary and impracticable to determine whether cyanic acid is a carbimide or the hydroxide of cyanogen, and whether prussic acid is a nitrile or an isonitrile. . . . It is scarcely necessary to add that the ideas which I have here developed are an application to the principles of chemical structure of the dynamic theory, the foundations of which were laid by Berthollet."

For this condition of dynamic equilibrium between isomers, no new name was proposed by Butlerow; but it may conveniently be referred to as DYNAMIC ISOMERISM, a name

which has the merit of summarizing—without adding anything to—the descriptive phrases which have been quoted above.

Tautomerism.

(a) In 1883 Baeyer 1 observed that isatin, like prussic acid, was capable of yielding two isomeric ethyl-derivatives. As only one form of the parent substance was known, Baeyer suggested that ethyl-pseudo-isatin was derived from a labile isomeride, which reverted to isatin whenever attempts were made to prepare it:

$$\begin{bmatrix} C_{\bullet}H_{\bullet} & CO \\ NH & NH \end{bmatrix} \rightarrow C_{\bullet}H_{\bullet} & CO \\ pseudo-Isatin. & Isatin. \\ \end{bmatrix}$$

In 1884 Zincke 2 made a converse observation to the effect that the phenylhydrazone derived from naphthaquinone was identical with the phenylazo-derivative of α-naphthol, with which it should have been isomeric.

C₆H₅·NH·N:C₁₀H₆:O C₆H₅·N:N·C₁₀H₆·OH
Phenylhydrazone Azo-compound Azo-compound.

All these observations could have been interpreted by Butlerow's theory of "dynamic equilibrium between isomers"; but this theory was hidden away under the title "Ueber Isodibutylen" and was not generally known. They were therefore discussed in 1885 on an independent basis by Laar in a paper "Ueber die Möglichkeit mehrerer Struktur-formeln für dieselbe chemische Verbindung". This paper attracted much wider attention on account of the provocative character of the title. The view was there advanced that the identical product obtained by the action of nitrous acid on phenol or of hydroxylamine on quinone, has not one but both of the constitutions shown in the scheme

$$ON \cdot C_6H_4 \cdot OH \rightleftharpoons HO \cdot N : C_6H_4 : O$$

For this purpose a hydrogen atom was supposed to oscillate

¹ Ber., 1883, 16, 2188.

² Ber., 1884, 17, 3030.

³ Ber., 1885, 18, 648.

incessantly between the two positions indicated in the formula

in a manner comparable with the vibrations that give rise to light. This phenomenon was described as tautomerism and was defined more strictly in the following year in a paper "Ueber die Hypothese der wechselnden Bindung", in which (in contrast to Butlerow's theory of interconvertible isomerides) it was postulated that quinone monoxime and nitrosophenol, or the two forms of prussic acid, were merely phases in the motion of a vibrating molecule, which could not be prepared as separate isomers, since the various formulæ that could be assigned to a tautomeric substance were defined as representing "not isomeric but identical substances". The phenomenon discovered by von Baeyer, that a substance, to which a perfectly definite structure could be assigned, might yield derivatives having a different structure, was described by Laar as pseudomerism.

Many of the substances cited by Laar as examples of tautomerism have been proved by further experience to be merely compounds which readily undergo isomeric change, but there is still a residue in reference to which precise information is still lacking. Thus nitrosophenol crystallizes from hot water in colourless needles, but from ether in large green plates, whilst the sodium salt crystallizes in red needles with 2H₂O. Since quinones are usually yellow in colour, whilst nitrosocompounds are often blue or green, we may infer that the red crystals of the sodium salt are quinonoid and that the green crystals actually consist of nitrosophenol and not of quinone monoxime.

The colourless needles are less easy to identify, but may perhaps be a dimeric form of nitrosophenol, since nitrosocompounds often lose their colour (like NO₂) on forming double molecules.¹ In the same way the colourless and yellow forms of *iso*nitrosomalonanilide ² may be formulated as follows:

No kinetic studies of isomeric change have been made, however, in either of these cases, and the conditions of equilibrium are still entirely unknown.

Isomeric change is relatively easy to detect when the isomers have different chemical properties (e.g. ketone and enol, nitro and ψ -nitro), but is much more difficult to demonstrate when both isomers are of similar chemical type. Ingold and Piggott ³ therefore propose to describe as functional tautomerism those cases in which a facile change of structure is accompanied by a change of chemical type, and as virtual tautomerism those cases in which there is no change of type. In order to demonstrate the existence of mobility of structure in cases of the latter group they propose three tests, ⁴ as follows:

(a) The Symmetry Test. The same product is obtained from both syntheses:

e.g.
(i)
$$C_0H_3:N_2Cl + NH_2:C_0H_4:CH_3 \rightarrow C_0H_3:N=N-NH:C_0H_4:CH_3$$

 $C_0H_3:NH_2 + ClN_3:C_0H_4:CH_3 \rightarrow C_0H_3:NH-N=N:C_0H_3:CH_3$

(b) The Fission Test. Degradation-products derived from two alternative structures are obtained by the fission of an apparently-homogeneous compound.⁵

¹ Bamberger and Seligman, Ber., 1903, 36, 689.

² M. A. Whiteley, J. Chem. Soc., 1903, 83, 34.

³ J. Chem. Soc., 1923, 123, 1469.

⁴ Cf. J. W. Baker, "Tautomerism" (Routledge, London, 1934), pp. 93 et seq.

⁵ Negative evidence, however, is inconclusive.

(c) The Substitution Test. When two isomers are obtained on replacing an atom of hydrogen by some other radical, it is assumed that the hydrogen is mobile; e.g. glucose gives two methyl-glucosides.

The value of these tests, however, depends in part on a knowledge of the mechanism of reaction. For instance, camphor yields an enolic benzoate when boiled with benzoyl chloride. This may be used as evidence that camphor can exist (or can at least react) in an enolic form,

$$C_{\textbf{a}}H_{\textbf{14}} \overset{\text{CH}_{\textbf{2}}}{\underset{\text{CO}}{|}} \rightleftharpoons C_{\textbf{a}}H_{\textbf{14}} \overset{\text{CH}}{\underset{\text{COH}}{|}}$$

This reaction can, however, also be interpreted as depend-

ing on the formation of an intermediate addition product

$$C_{0}H_{1} \leftarrow \begin{array}{c} CH_{0} \\ C=0 \end{array} \xrightarrow{\pm \begin{array}{c} C_{0}H_{1} \cdot COCl \\ C=0 \end{array}} C_{0}H_{1} \leftarrow \begin{array}{c} CH_{0} \\ C=0 \end{array} \xrightarrow{-HCl} CH_{0} \leftarrow CCCCC_{0}H_{0}$$

It therefore provides no real proof of the existence of the hypothetical enol, although this may be postulated on other grounds.

Desmotropy and Prototropy.

Laar applied his theory of "tautomerism" to a large group of labile compounds, the interconversion of which involves the migration of an atom of hydrogen, and a rearrangement of single and double bonds, as in ethyl acetoacetate, to which an enolic formula had been assigned by Geuther 1 and a ketonic formula by Frankland and Duppa 2 in order to explain its diverse reactions. It was not long, however, before compounds of this type were isolated in isomeric forms and were thus shown conclusively not to be

¹ Tahresber., 1863, 323.

² Proc. Roy. Soc., 1865, 14, 458; J. Chem. Soc., 1866, 19, 395; 1867, 20, 102; Phil. Trans., 1866, 156, 37.

"tautomeric" in the sense of the original definitions. Thus Claisen in 1893 isolated two forms of acetyldibenzoylmethane

$$\begin{array}{ccc} \text{CH}_{s}\text{-CO-CH}(\text{CO-C}_{e}\text{H}_{s})_{2}, & \text{CH}_{s}\text{-C(OH):C(CO-C}_{e}\text{H}_{s})_{2} \\ \text{Ketonic form.} & \text{Enolic form.} \\ \text{(m.p. 107-110°)} & \text{(m.p. 80-85°)} \end{array}$$

and Knorr in 1899 prepared no less than five forms of ethyl diacetylsuccinate.

In the classical example of ethylacetoacetate, Knorr 3 isolated the ketonic form as a solid melting at -39° , by cooling a solution of the ester in ether, alcohol or light petroleum to -80° , whilst the enolic form was obtained by precipitating sodium chloride by the action of gaseous hydrogen chloride from a suspension of the sodium salt at the same low temperature; it was then separated as a glassy solid by evaporating the solvent, and was distilled under reduced pressure at 33° , without undergoing appreciable change. By measurements of refractive indices the proportions of the two isomers at equilibrium were estimated as follows:

- (i) Ketone, $n_D^{10} = 1.4225$
- (ii) Enol, $n_D^{10} = 1.4480$
- (iii) Equilibrium mixture, $n_D^{10} = 1.4232$: Enol = 2%.

In a similar manner Kurt Meyer⁴ separated the two forms of the ester by "aseptic distillation", i.e. by distillation under reduced pressure in silica vessels, which (unlike glass) did not catalyse the interconversion of the two isomerides. He also showed ⁵ that, whilst the enolic form was acted on immediately by bromine, the ketonic form was inert; the proportion of enol could therefore be estimated by adding bromine, removing the excess by the addition

¹ Annalen, 1893, 277, 191; 1896, 291, 73.

² Annalen, 1899, 306, 332-93.

³ Ber., 1911, 44, 1138. ⁴ Ber., 1920, 53, 1410.

⁵ Annalen, 1911, 380, 212.

of phenol and then titrating the iodine set free from potassium iodide by the brominated enol.

(i)
$$CH_3$$
— C — CH — $COOEt + Br_2$ \longrightarrow CH_3 — C — CH — $COOEt$

OH

OB

(ii) CH₈—C—CH—COOEt + 2HI
$$\rightarrow$$
 CH₈—C—CH—COOEt + O Br O H

 $HBr + I_{\bullet}$

The compound produced by reaction (i) contains a "positive bromine atom" (cf. p. 222). In this way the proportion of enol in different solvents was estimated to be as follows:

Water		0.4%	Ethyl acetate	12.9%
Acetic acid .		5.74%	Benzene	16.2%
Methyl alcohol			Ether	
Acetone			Carbon disulphide	32.4%
Chloroform .			Hexane	
Nitrobenzene			No solvent	7.7%
Ethyl alcohol	•	12.0%		

Since Laar's theory is obviously untrue in the cases to which he applied it, new terms and new definitions are needed in order to describe (a) the theory of reversible isomeric change, which provides a correct explanation of the dual reactivity to which he had directed attention, and (b) the particular type of isomeric change which is characteristic of the compounds cited in his paper, namely that which involves the migration of a mobile hydrogen atom and a rearrangement of single and double bonds.

- (a) The first phenomenon is adequately covered by the term dynamic isomerism, which was introduced by Lowry in 1899 in order to describe all cases of reversible isomeric change, whether provoked by the deliberate addition of a catalyst or proceeding apparently spontaneously (but perhaps under the influence of casual impurities) in material of a normal degree of purity.
 - (b) The special type of isomerism referred to above is

¹ Trans. Chem. Soc., 1899, 75, 211.

still described by many writers as "tautomerism", since the compounds which Laar selected to illustrate his theory were all of this type. Thus in a translation of Schmidt's "Text-book of Organic Chemistry", dated 1926, we read:

"A substance is tautomeric when it forms two series of derivatives. These are derived from two parent structures which differ only in the position of a hydrogen atom and of one or more double bonds."

In the same way Kurt Meyer 1 suggests that

"Substances are tautomeric, if they form two series of derivatives which are deduced from two isomeric formulæ; these formulæ differ from one another in the position of a hydrogen atom, and of one or more double bonds."

A much more satisfactory proposal was made, as long ago as 1887, by Jacobsen,² who was one of the first to repudiate both the theory of Laar and the name by which it was described:

"The word 'tautomerism' is based on Laar's view, which (I believe) is not shared by most chemists, that the molecules of compounds, whose chemical behaviour is represented by two structural formulæ differing in the point of attachment of a hydrogen atom, never assume a definite constitution, but exist in a constant state of oscillatory change. The majority of chemists would explain the observations in question in this way, that the known forms of such compounds are to be represented by a definite grouping of atoms which in certain reactions passes over into an isomeric grouping by a rearrangement of bonds consequent upon the displacement of a hydrogen atom."

This "rearrangement of bonds consequent upon the displacement of a hydrogen atom" was described by Jacobsen as desmotropy ($\delta \epsilon \sigma \mu \rho \varsigma$, a bond, $\tau \varrho \delta \pi \epsilon \nu$, to turn) and by Michael 3 as merotropy; but the former term was rendered ambiguous (and therefore useless for the purpose of

¹ Annalen, 1913, 398, 63.

² Ber., 1887, 20, 1732 footnote; 1888, 21, 2628 footnote.

^{3 &}quot;Ueber Desmotropie und Merotropie", Annalen, 1908, 363, 20 et seq.

precise definition) by Hantzsch and Hermann, who rejected Laar's theory but proposed (i) to use his title as a general description of compounds which could react in two or more different ways, and (ii) to divert Jacobsen's term from its original meaning, in order to describe those cases of pseudotautomerism in which isomeric forms could be crystallized out. Since Jacobsen's "bond-shifting" is no longer available as an unequivocal description of the migration of a proton and the rearrangement of valency-electrons which accompanies it, the other and more characteristic aspect of this dual process has been made the basis of a new description of the phenomenon as "proton-shifting" or prototropy. Prototropy must be regarded as a particular case of the general phenomenon of ionotropy or "ion shifting".

On the other hand, since no new term has been proposed in order to describe the phenomenon of dual reactivity, widespread support has been given to the definition provided by an anonymous writer in the Oxford Dictionary, who defines the phenomenon of tautomerism in the following terms:

"This term is applied to the property exhibited by certain compounds of behaving in different reactions as if they possessed two or more different constitutions; that is, as if the atoms of the same compound or group were arranged in two or more different ways, expressible by different structural formulæ."

This definition has the following merits:

(i) It abolishes finally the distinction, upon which Laar insisted so strongly, between "tautomerism" and "pseudomerism". (ii) It defines tautomerism as a purely chemical phenomenon, and thus distinguishes it sharply from the physico-chemical phenomenon of dynamic isomerism. (iii) It abolishes the former limitation of the term to prototropic

¹ Ber., 1887, 20, 2802.

² Lowry, J. Chem. Soc., 1923, 123, 828.

³ Ingold, Annual Reports of Chemical Society, 1927, p. 106.

compounds and extends it to mobile structures of other types.

Examples of tautomerism as defined by Laar are found in molecules and ions capable of resonance (p. 393), as, for example, methyl azide:

$$CH_3-N=\stackrel{+}{N}=\stackrel{-}{N}$$
 : $CH_3-\stackrel{-}{N}=N$: $CH_3-\stackrel{-}{N}-\stackrel{+}{N}=N$

Mathematical analysis indicates that, in molecules possessing resonance structures, valency electrons may be displaced without any rearrangement of atomic nuclei, and though a molecule may react chemically as if it possessed either structure (A) or structure (B), it is impossible to adduce physical evidence for the existence of either of these "unperturbed" forms. Ingold has described the stable form as a mesomeric state, and has introduced the symbolism of (C) above, in which looped signs (without arrows) are inserted to indicate the wide locality within which the valency electrons may be found.

Dual Reactivity of Dynamic Isomerides.

Any substance which exhibits dynamic isomerism under the conditions which prevail during a given chemical reaction must necessarily exhibit dual reactivity also. Thus, although the action of bromine on ethyl acetoacetate may be limited to the enolic form, the pure β -diketone must necessarily react with the halogen if time and opportunity are given for isomeric change to take place. Under suitable conditions, therefore, a compound may be converted quantitatively into a derivative of an isomeride which is present only in traces in the equilibrium-mixture, as in the action of bromine on acetone, the velocity of which was shown by Lapworth to be independent of the concentration of bromine, but directly proportional to the concentration of the acid catalyst by which the ketone was brought into equilibrium with a trace of the enolic isomeride.

The converse of the preceding proposition, however, is

¹ J. Chem. Soc., 1933, 1120; Chemical Reviews, 1934, 15, 250.

not necessarily true, since dual reactivity may appear under conditions which do not give rise to dynamic isomerism. The contrast which thus exists between dual reactivity and reversible isomeric change may be illustrated by the behaviour of α - and α' -chlorocamphor and bromocamphor on nitration. These pairs of isomers are perfectly stable in the solid state or in neutral alcoholic solutions; but they undergo reversible isomeric change in alkaline solutions 1 :

$$C_{\mathfrak{d}}H_{1\mathfrak{d}}$$
 $C_{\mathfrak{d}}H_{1\mathfrak{d}}$
 $C_{\mathfrak{d}}H_{1\mathfrak{d}}$

On nitration, the two bromocamphors yield identical mixtures of $\alpha\alpha'$ - and $\alpha'\alpha$ -bromonitrocamphor, and therefore exhibit the phenomenon of dual reactivity as described in the Oxford definition of tautomerism. Similarly, on bromination, the two chlorocamphors give identical mixtures of the stereoisomeric $\alpha\alpha'$ - and $\alpha'\alpha$ -chlorobromocamphors. On nitration, however, they yield the isomeric $\alpha\alpha'$ - and $\alpha'\alpha$ -chloronitrocamphors as pure products, without giving any indication of alteration of structure under the influence of the nitric acid used as a reagent.

The Mechanism of Prototropic Change.

(a) Mutarotation. One of the most effective methods of studying the more facile types of chemical change is by observations of the rotatory power of optically-active compounds of an appropriate type. Thus, as long ago as 1846 spontaneous alterations of rotatory power were observed by Dubrunfaut 4 in freshly-prepared solutions of glucose in water, and these changes were shown in subsequent years

¹ Kipping, Proc. Chem. Soc., 1905, 125; Lowry, J. Chem. Soc., 1906, 89, 1033; Lowry and Steele, ibid., 1915, 107, 1382; Lowry, Steele and Burgess, ibid., 1922, 121, 633.

² Lowry, Steele and Burgess, J. Chem. Soc., 1922, 121, 635.

³ Lowry and Steele, J. Chem. Soc., 1915, 107, p. 1389.

⁴ Compt. rend., 1846, 23, 38

to be a general characteristic of the reducing sugars. Dubrunfaut himself (at a time when structural formulæ had not yet been developed) regarded these alterations of rotatory power as an effect of changes in the molecular structure of the sugar, whilst Landolt in 1879 1 attributed them to a progressive conversion of the large molecular aggregates of the crystal into the smaller units of the chemical molecule. As an alternative, Emil Fischer,2 who had observed similar changes of rotatory power during the reversible hydrolysis of gluconic lactone, suggested that the sugars, like the lactones prepared from them, became hydrated in freshly-prepared solutions in water:

$$C_0H_{10}O_0 + H_2O \rightleftharpoons C_0H_{12}O_7$$

Gluconic lactone. Gluconic acid.
 $C_0H_{12}O_0 + H_2O \rightleftharpoons C_0H_{14}O_7$
Glucose. Glucose hydrate.

This uncertainty as to the origin of the phenomenon did not exist in the case of nitrocamphor, which was found in 1899 to exhibit changes of rotatory power in freshly-prepared solutions in a wide range of anhydrous solvents.³ These changes of rotatory power were described as *muta-rotation* ⁴ and were attributed to a *reversible isomeric change* or dynamic isomerism, as represented by the equation:

$$C_{\mathfrak{g}}H_{14}$$
 $C_{\mathfrak{g}}H_{14}$
 $C_{\mathfrak{g}}H$
 $C_{\mathfrak{g}$

(b) Optical Inversion as a Cause of Mutarotation. By analogy with nitrocamphor, the mutarotation of the sugars was attributed to reversible isomeric change 5 ; but the structure of the isomerides was not obvious, until Simon in 1901 6 suggested that α - and β -glucose, which had

^{1 &}quot; Optisches Drehungsvermögen."

² Ber., 1890, 23, 2625.

³ Lowry, J. Chem. Soc., 1899, 75, 211.

⁴ Loc. cit., p. 213. ⁵ Lowry, J. Chem. Soc., 1899, 75, 211.

⁶ Compt. rend., 1901, 132, 487.

been isolated by Tanret in 1896, were the lower homologues of the α - and β -methylglucosides, since they had the same average rotatory power

α-Methylglucoside + 157°
$$-32$$
 α-Glucose + 105° -32 γ-Glucose + 22

Mean + $62 \cdot 5$ ° $-63 \cdot 5$ °

Since a cyclic formula had already been attributed to the glucosides, it followed that the mutarotation of α - and β -glucose was primarily due to the interconversion of two stereoisomeric forms of a cyclic sugar ²

This conclusion was confirmed by the experiments of E. F. Armstrong, who showed that, when hydrolysed with an appropriate enzyme, the α - and β -methylglucosides gave rise to glucoses which exhibited mutarotation in the same direction as α - and β -glucose. Since, however, the optically-active >CHOH groups of the sugars are remarkably stable, and can only be inverted in exceptional cases, it was unlikely that this could take place spontaneously whilst mutarotation was proceeding in cold aqueous solutions. It was therefore suggested α that the mutarotation of a hexose in presence of water could be attributed to the reversible formation of an aldehydrol, whereby the hydrated α - and β - (pyranose)

¹ Bull. Soc. Chim., 1896, 15, 195 and 349. The β -sugar was prepared by evaporating an aqueous solution at 100°.

² In the formulæ now cited, the 5-ring (furanose) structure, which was at that time assigned to the glucosides, has been replaced by a 6-ring (pyranose) structure.

³ Lowry, J. Chem. Soc., 1904, 85, 1551.

sugars could be brought into equilibrium as follows:

$$\begin{array}{c|ccccc} H & OH & & HO & H \\ \hline & C & CH(OH)_2 & & C \\ H & & & & HO & H \\ O & [CHOH]_3 & \rightleftharpoons & [CHOH]_3 & & HO & |CHOH]_3 \\ HO & & & & & HO & |CHOH]_3 & & HO & |CHOH]_3 \\ CH & & & & & & CHOH & & CH \\ CH_2OH & & & & & CH_2OH & & CH_2OH \\ \alpha\text{-Glucose}. & & & & Aldehydrol. & & \beta\text{-Glucose}. \end{array}$$

Similarly in the absence of water the α - and β -pyranose sugars may be brought into equilibrium through the reversible formation of the aldehydic form of the hexose, e.g. during the mutarotation of methylated and acetylated sugars in non-aqueous solvents.¹

Unlike the α - and β -sugars, the aldehydic sugars, which are shown as intermediate products in these schemes, are not known in the free state, but stable derivatives have been isolated by methylating,² acetylating ³ or benzoylating ⁴ the thioacetals, $HO\cdot[CHOH]_{\$}\cdot C(SEt)_{2}$, and then eliminating the mercaptyl groups. A mere trace of the open-chain aldehyde in equilibrium with a pyranose or furanose sugar would suffice, however, to bring the cyclic forms of the sugar into equilibrium with one another. In the same way, a minute trace of enol in equilibrium with an optically-active carboxylic acid will suffice to bring about the optical inversion of the asymmetric carbon atom

¹ Purdie and Irvine, *J. Chem. Soc.*, 1904, **85**, 1049; Fischer and Delbrück, *Ber.*, 1909, **42**, 2778; Irvine and Scott, *J. Chem. Soc.*, 1913, **103**, 575.

² Levene, J. Biol. Chem., 1926, 69, 175; 1927, 74, 695.

³ Wolfrom, J.A.C.S., 1929, **51**, 2188 et seq.

⁴ Brigl and Mühlschlegel, Ber., 1930, 63, 1551.

310 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

In this way a simple mechanism can be found to account for the conversion of d- and l-tartaric acid into meso- and then into racemic or dl-tartaric acid under the influence of concentrated alkali, and for the conversion of camphoric acid into isocamphoric acid under the influence of concentrated hydrochloric acid.

The correctness of the mechanism thus suggested is shown

by the fact that in camphoric acid only the COOH

group is inverted, and not the COOH group in which

the mobile hydrogen has been replaced by methyl, so that complete racemization (like that of tartaric acid) does not occur. The mutarotation of the sugars, which depends on

group, requires a different mechanism, but the only essential points of difference are

- (i) That the carbonyl-form of the sugar is the labile instead of the stable isomer.
- (ii) That the carbonyl-compound is converted reversibly into a cyclic isomeride containing a 6-atom or 5-atom ring, instead of into an enol containing only a double bond or "2-atom ring".1
 - (iii) That the labile isomeride is formed much more
- ¹ This process was described by Jacobsen and Stelzner as "oxocyclo-desmotropy" and was cited by Baker, Ingold and Thorpe as an example of "Ring-chain Tautomerism" (J. Chem. Soc., 1924, 125, 269).

readily than in the carboxylic acids cited above, since isomeric change proceeds rapidly in cold aqueous solutions, instead of requiring the use of hot concentrated acid or alkali as a catalyst.

(c) Mutarotation is not Spontaneous in Anhydrous Solvents. The discovery that solutions of nitrocamphor undergo mutarotation in solvents which are not miscible with water was followed immediately 1 by the further discovery that the changes of rotatory power are not spontaneous, but depend on the presence of a catalyst. This conclusion was based on observations of the rotatory power of a solution

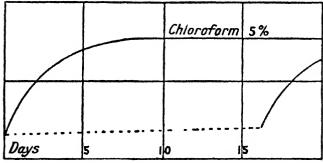


FIGURE VII.—Change in Rotatory Power of Nitrocamphor in Solution in Chloroform.²

of nitrocamphor in chloroform (Fig. VII), which fell from $[\alpha]_D - 27^\circ$ to a steady value of -14° , in the course of 8 days. At the end of 16 days the remainder of the solution was found to be almost unchanged; but mutarotation was again produced by transferring the solution from the graduated flask (of hard glass?) in which it had been prepared to a polarimeter tube of soft glass in order to determine its rotatory power.

Later,3 the unexplained arrest of mutarotation in chloro-

¹ Lowry, J. Chem. Soc., 1899, 75, 211.

² By permission of Professor Lowry from an original paper by him in "Chemical Reviews," 1927, published by The Williams and Wilkins Co., Baltimore, U.S.A.

⁸ Lowry and Magson, J. Chem. Soc., 1908, 93, 119.

form was traced to the anticatalytic properties of carbonyl chloride, formed by oxidation of the solvent:

$$CHCl_3 + O \longrightarrow COCl_2 + HCl.$$

Its activity as an anti-catalyst was attributed to the irreversible elimination of basic impurities, e.g. by converting ammonia into carbamide

$$COCl_2 + 4NH_3 = CO(NH_2)_2 + 2NH_4Cl.$$

The addition of traces of carbonyl chloride was then found to give rise to similar "arrests" of mutarotation, in benzene or ether, provided that the solutions were kept in silica flasks, and transferred to glass polarimeter tubes only when it was necessary to take out a sample in order to observe the progress of the change.

Similar evidence cannot readily be obtained in reference to the mutarotation of natural sugars, such as glucose, galactose and lactose, since they are difficult to purify completely from water-soluble impurities and are insoluble in the most promising solvents. Arrests of mutarotation have, however, been observed in solutions of tetramethylglucose in chloroform, benzene, ethyl acetate and pyridine, as well as in solutions of tetra-acetylglucose in ethyl acetate.²

(d) Catalysis of Mutarotation by Amphoteric Solvents. The mutarotation of aqueous solutions of reducing sugars is catalysed by acids 3 and still more effectively by alkalis. 4 In neutral or feebly-acid solutions the velocity of mutarotation falls to a minimum; but it cannot be arrested by adding a trace of acid to neutralize alkaline impurities. 5 Water is, therefore, evidently a catalyst, apart from any traces of acid or alkali which may be present. Its catalytic activity is in fact 10 times greater than that of the hydrogen and hydroxyl ions which it contains. Hudson therefore

¹ Lowry and Richards, J. Chem. Soc., 1925, 127, 1385.

² Lowry and Owen, Proc. Roy. Soc., 1928, A 119, 505.

³ Erdmann, "Dissertatio de sacharo lactico et amylaceo": cited in Jahresbericht für 1855, p. 671.

⁴ Urech, Ber., 1882, 15, 2132.

⁵ Lowry, J. Chem. Soc., 1903, 83, 1314; Chemical Reviews, 1927, 4, 231.

expressed the velocity of mutarotation by the equation

$$k = 0.0096 + 0.258 \text{ [H]} + 9750 \text{ [OH]}$$

where the constant term represents the catalytic activity of the pure solvent. Mutarotation in aqueous solutions, however, is catalysed by acids and bases in the sense of the extended definition discussed in Chapter XII. Any ion or molecule which will give or accept a proton will therefore accelerate the process. Thus the following values have been determined for the catalytic coefficients of acid and basic molecules and ions as deduced from experiments on the mutarotation of glucose at 20°.

TABLE I

Catalytic coefficients for Glucose at 20°

Molecules.

[HCl] about 0.5

[OH] about 8000

[HAc] = 0.0065

[OH₃] or [H] = 0.36

[OH₄] = 0.00026

[Ac]

= 0.069

A much longer series of catalytic coefficients, based upon the results of dilatometric measurements, has been given by Brönsted and Guggenheim, who have shown that the value of the coefficient k is related to the dissociation constant K of the acid or base according to the equation

$$k = GK^{x}$$

where G is an arbitrary constant, and the exponent x is about 0.4 for basic and 0.2 for acid catalysis, in presence of a large excess of water.

In the absence of water, neither acids nor bases appear to possess independent catalytic properties. Thus pyridine, which was regarded by Behrend and Roth 2 as an efficient catalyst (comparable with water) for the mutarotation of glucose, does not act as a catalyst when reason-

¹ J.A.C.S., 1927, 49, 2554.

² Behrend and Roth, Annalen, 1904, 331, 359.

ably pure and dry, since the mutarotation of tetramethyl-glucose can be "arrested" over a period of 2 or 3 hours in this solvent, after which the solution gradually becomes green and mutarotation proceeds with progressively increasing speed (Fig. VIII). Conversely, cresol appears to possess no catalytic activity when dry; but mixtures of pyridine and cresol (in which the acid and basic properties are presumably less pronounced than in the two pure solvents) promote catalysis, since the maximum velocity of mutarotation of tetramethylglucose in a mixture of I part of pyridine and 2 parts of cresol is about 20 times greater than in water 1 (Fig. IX). Complete catalytic activity is therefore characteristic of amphoteric solvents, such as water or a mixture of pyridine and cresol, which possess both acid and basic properties, and can therefore both give and accept a proton.

(e) Electron-migration in Prototropic Change. Since prototropic change is catalysed by acids and bases, i.e. by proton donors and acceptors, it is clear that this process depends on a migration of a proton, H, and not of a neutral hydrogen atom. In the simple case of a ketoenolic change, we may note that the removal of proton from either isomer gives rise to an identical ion, which may be indicated either by writing the double bond in the ion in the "activated" or semi-polar form, thus:

or, by considering that the common ion exists in a mesomeric state, as

$$\left[-\frac{1}{2} - \frac{1}{2} - \frac{1}{2} \right]^{-1}$$

¹ Lowry and Faulkner, J. Chem. Soc., 1925, 127, 2883.

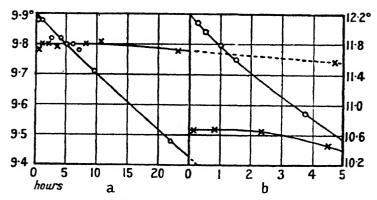


FIGURE VIII.1—(a) Mutarotation of Tetramethylglucose in "Dry" and "Wet" Ethyl Acetate. x, dry; o, wet.

(b) Mutarotation of Tetramethylglucose and Glucose in "Dry" Pyridine. o, glucose; x, tetramethylglucose.

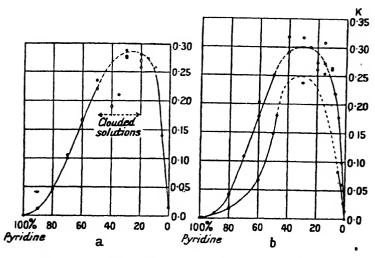


FIGURE IX.1—(a) Mutarotation of Glucose in Pyridine and Water.
(b) Mutarotation of Tetramethylglucose (i) in Pyridine and Water,
(ii) in pyridine and cresol.

¹ Reproduced, by permission, from Chemical Reviews, 1927, 4.

316 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

The dynamic isomerism of the system may therefore be attributed to a reversible separation of a proton, coupled with a movement of electrons in the resulting ion. The effect of this migration is to transfer a negative charge from one possible point of attachment of the proton to the other, and thus to make provision for neutralizing its positive charge.

The whole process is obviously electrical in character, depending on ionization and not on the formation of free radicals. It may be compared with an electrolysis, in which the acid and base act as negative and positive poles respectively, whilst the organic molecule takes the place of the electrolyte. Thus if, with Brönsted and Guggenheim, we write the two forms of the prototropic compound as HS and SH, the isomeric change in an amphoteric solvent may be represented by the reversible equation

$$B + HS + HA \rightleftharpoons BH + SH + \overline{A}$$

where B is the basic and HA is the acid catalyst. It is, however, essential that a negative charge shall be able to travel through the radical S from one terminal point to the other, in order to neutralize the negative and positive charges which are left behind by the removal and addition of a proton; and this property is in fact the principal characteristic, not only of the unsaturated and cyclic systems of prototropic compounds, but also of conjugated systems (Chapter XV), in which a similar migration must be supposed to occur along the whole length of the conjugated chain,² e.g. in order to account for the neutralization of opposite charges at the terminal points of the chain, when addition of an ionized reagent takes place in the 1:4 or 1:6 positions, as in the addition of HCN to an unsaturated ketone.

Valency Changes in Tautomerism.

In his original paper on tautomerism Laar classified isomeric changes as "dyad", "triad", "pentad" changes,

¹ J.A.C.S., 1927, 49, 2554. ² Lowry, Nature, 1925, 115, 376.

etc., according to the number of atoms separating the initial and final positions of the migrating hydrogen atom, but did not further differentiate between them. Lapworth, in 1898, pointed out that when the system involved an odd number of atoms the migration of the hydrogen did not alter the valency of any atom, but that when an even number of atoms were involved, as in the simple dyad system $H-CN \rightleftharpoons CN-H$, then the valency of one of the terminal atoms changed by two units. Dyad or tetrad systems must therefore contain a terminal atom of variable valency number, such as nitrogen, oxygen or sulphur.

The increase of valency is, however, made up of one covalency and one electrovalency, since one of the isomers has the structure of an "onium" compound, e.g. :C=N-H (p. 108), and the electronic migration in the organic molecule occurs by exactly the same process as in the triad keto-enol system previously exemplified (p. 314).

"Spontaneous" Isomeric Change in Liquid or Gaseous Prussic Acid. In his classical paper on "iso-Dibutylene", Butlerow suggested that liquid prussic acid could undergo reversible isomeric change without the addition of a catalyst:

$$H \cdot C \equiv N \rightleftharpoons H \cdot \stackrel{+}{N} \equiv \stackrel{-}{C}$$
:

Subsequent experiments 2 have shown that the liquid acid is almost pure hydrogen cyanide, containing not more than $\frac{1}{2}$ % of hydrogen isocyanide. In this respect prussic acid may be compared with acetone, which is generally believed to exist almost exclusively in the ketonic form. This case is therefore not far removed from the "pseudomerism" of isatin, where the equilibrium is so one-sided that Baeyer regarded the less stable isomeride as non-existent.

Butlerow's suggestion that isomeric change may pro-

¹ J. Chem. Soc., 1898, 73, 457,

² See for instance Usherwood, J. Chem. Soc., 1922, 121, 1604; Dadieu, Ber., 1931, 64, 358; Lowry and Henderson, Proc. Roy. Soc., 1932, A 136, 471.

ceed in the absence of a catalyst is certainly not true of acetone, where the velocity of iodination depends primarily on the speed with which the ketone can be enolized in presence of an appropriate catalyst; and it is also not true of other $\alpha\gamma$ -migrations, where the wandering of a hydrogen atom between the α - and γ -positions can be arrested (as described above) either by rigorous purification, or by the addition of an anti-catalyst. No analogous evidence has, however, been found in the case of prussic acid. Thus Kurt Meyer and Hopff 1 were unable to effect any separation of isomers by the method of "aseptic distillation"; and, as long ago as 1869, Gautier 2 failed to effect any separation by the fractional crystallization of not less than 600 grams of the liquid acid.

These negative results are not conclusive; but the hypothesis of spontaneous isomeric change is certainly more plausible in $\alpha\beta$ - than in $\alpha\gamma$ -migrations, in view of the fact that the two forms of prussic acid, from which the alkyl cyanides and isocyanides are derived, yield ions which are not isomeric but identical, thus:

$$CH_{3} \cdot C = N \leftarrow H \cdot C = N \rightleftharpoons H + \overline{C} = N \rightleftharpoons H \cdot \overline{N} = \overline{C} \longrightarrow CH_{3} \cdot \overline{N} = \overline{C},$$
or
$$H: C:::N: \rightleftharpoons H + :\overline{C}:::N: \rightleftharpoons H: \overline{N}:::\overline{C}:$$

Since the two hydrides shown in this equation could be converted into one another by mere ionization, and since prussic acid is itself an excellent ionizing solvent for other electrolytes, it is possible that isomeric change in this particular system might take place in the pure liquid, in the absence of any foreign substance to act as a catalyst.

Anionotropic Change.3

Isomeric change does not necessarily involve the migration of a proton, though undoubtedly, on account of the great ease of release of hydrogen from acidic substances,

¹Ber., 1921, 54, 1709. ² Ann. Chim. Phys., 1869, 17, 103.

³ For numerous examples, see J. W. Baker's "Tautomerism", 1934 (Routledge, London).

prototropic change is much more facile than any other type of isomerization.

Just as a base can bring about the release of proton and catalyse chemical change in a prototropic system, so an acid can frequently bring about the release of an anion, such as the hydroxyl ion OH⁻, and initiate an anionotropic change, e.g.—

Anionotropic change is by no means confined to isomerization dependent upon the mobility of hydroxyl, but can occur whenever a catalysing solvent can promote the release of a negatively charged ion from an organic molecule. For example, Burton 2 has shown that the mobility of compounds of the general formulæ

R—CH(X)—CH=CH—R' \rightleftharpoons R—CH=CH—CH(X)—R' follows the order of anionic stability of X (e.g. Br>OAc>OH) and, further, has demonstrated that a complete ionic dissociation occurs during the course of the isomerization. Thus the alcohols (I) and (II) are obtainable separately, and each can be esterified without a change of structure.

A conversion from the phenyl-allyl ester to the cinnamyl ester can be effected by boiling for six hours in acetic anhydride. Isomerization is slower in benzonitrile or chlorobenzene, and is almost negligible in boiling xylene. The corresponding bromohydrocarbons are so mobile that isomerization takes place immediately a solution of phenyl-

¹ J. Chem. Soc., 1928, 906. ² Ibid., 1928, 1653; 1929, 455.

allyl alcohol in glacial acetic acid is treated with hydrogen bromide, only the cinnamyl bromide being capable of isolation. The first stage in the reaction is a complete ionic dissociation, since it is possible to replace one anion by another. For example, when phenyl-allyl p-nitrobenzoate is heated in acetic anhydride containing dissolved tetramethylammonium acetate (i.e. acetate ions) it yields cinnamyl p-nitrobenzoate plus a large proportion of cinnamyl acetate. This replacement is, of course, similar to the replacement of chlorine by hydroxyl in the reaction of olefines with chlorine-water (p. 175).

Replacement of one anion by another occurs whenever a pseudo-base forms a salt with an acid—a chemical change exactly analogous to the formation of a salt from a pseudo-acid and a metallic kation. The slow occurrence of both these neutralizations can be demonstrated by conductivity measurements, which prove conclusively that ionization must occur during the isomerization.

Migration of Hydrocarbon Radicals.

It has already been shown that both alkyl and aryl groups can give rise to both positive and negative ions, and hence it is not surprising that, under suitable conditions, hydrocarbon radicals should be able to migrate during an isomerization. Further, if one contrasts the ease of ionization of hydrogen or hydroxyl with the great difficulty of ionization of a hydrocarbon group then one would anticipate that the migration of a hydrocarbon group would never be facile. This is indeed the case, since alkyl and aryl radicals can only be made to migrate by heating their compounds to high temperatures.

For example, whilst hydrocyanic acid will isomerize spontaneously, the alkyl isocyanides will only isomerize to cyanides when their vapours are heated to a high temperature.

$$C_{\mathfrak{s}}H_{\mathfrak{s}}-\overset{+}{N}=\overset{-}{C}\longrightarrow C_{\mathfrak{s}}H_{\mathfrak{s}}-\overset{-}{C}=N.$$

¹ Nef, Annalen, 1894, 280, 296.

Similarly, the esters of thiocarbonic acid will only isomerize when heated above 300° C.1

$$Ar$$
— O C == S \rightarrow Ar — O C — S — Ar

The physical conditions under which aryl radicals will migrate have been investigated in detail by Chapman both for imino-aryl ethers and for substituted amidines.

(i) Imino-aryl Ethers. Benzanilide can be converted into the imide chloride by the action of phosphorus pentachloride, and the product, on treatment with sodium phenate, yields N-phenyl-benziminophenyl ether (I), a stable crystalline solid which on heating to 270–300° C. is smoothly converted into benzoyldiphenylamine (II).²

The progress of the isomerization can be followed by determination, from time to time during the change, of the melting-point of the mixture of (I) and (II). It was found to be strictly unimolecular, and must therefore occur by dissociation of (I) into either two ions or two radicals, which must reunite after a redistribution of valency bonds has occurred. Since a mixture of two different imino-aryl ethers gives on heating a mixture of only two, and not four, acylated amines it follows that the intermediate dissociation products must have too brief an existence to separate far enough to intermingle.

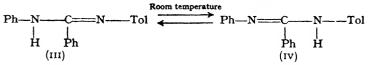
In order to test whether free ions are produced during the molecular rearrangement, the electrical conductivity of the reacting molten system was taken. This was found

¹ Schönberg and Vargha, Ber., 1930, 63, 178.

³ J. Chem. Soc., 1925, 1992; 1927, 1740.

to be low below 200° C., but increased simultaneously with the commencement of isomeric change. The conductivity rise was at first rapid, but the increase gradually became slower, and finally the conductivity remained constant at the value required for the pure end-product (II).

(ii) Amidines. A definite case of reversible isomerization is provided by substituted amidines. Von Pechmann in 1895 showed that the two compounds (III) and (IV) formed a mixture in tautomeric equilibrium, it being impossible to prove the exact location of the hydrogen atom:



and in 1929 Chapman 2 was able to demonstrate that the corresponding phenyl derivatives (V) and (VI), though easily separable and stable at room temperature, gave, on heating, an identical equilibrium mixture.

As in the case of the imino-aryl ethers, the isomerization cannot involve the production of ions having independent existence for any appreciable duration of time. In these reactions the migrating phenyl group replaces the mobile proton of the parent amide or amidine, and hence one would anticipate that the phenyl group must migrate in the form of a kation. It is impossible to decide this conclusively, however, by mere inspection of formulæ, since these can be written so as to indicate equally well either kationotropic or anionotropic change. Chapman considers that the phenyl group most probably migrates as an anion in both reactions since (a) the migration in the imino-aryl ethers occurs from oxygen to the less negative element nitrogen,

¹ Ber., 1895, 28, 869.

² J. Chem. Soc., 1929, 2133; 1930, 2458.

and (b) substituent groups (R) in a whole series of ethers (or amidines) of the general type R—O—C(Ph)=N—Ph facilitate the isomerization in the order nitrophenyl>chlorophenyl>phenyl>methyl, i.e. in the order of stability of their anions.¹ He has suggested, however, that the migrating phenyl group carries with it its bonding electron pair, but never moves from close juxtaposition of the other radical as a completely dissociated ion (cf. p. 367).

¹ J. Chem. Soc., 1927, 1744; 1930, 2464.

CHAPTER XIV

MOLECULAR REARRANGEMENT¹

Molecular Rearrangement and Isomeric Change.

In the preceding chapter, a description has been given of a very interesting group of isomeric changes, namely, ionotropic changes, which depend on the migration of an anion or a kation. These migrations are accompanied by the wandering of a double bond (or sometimes by the formation or rupture of a ring), as a result of which the charges carried by the migrating ion are neutralized by the migration of electrons within the molecule. The present chapter deals in a more general way with changes of molecular structure. These molecular rearrangements often take the form of an isomeric change, but in other cases the linkages between atoms are displaced during a process of substitution.

- (a) The simplest type of molecular rearrangement is the Walden inversion, in which an asymmetric carbon atom undergoes an optical inversion, or reversal of configuration, during the replacement of one radical by another; on reverse substitution therefore an isomeric change, $d \rightleftharpoons l$, is often found to have taken place. The final effect is identical with that which would be produced by an interchange of two radicals on the same carbon atom in one of the processes formulated above; but this is not necessarily the mechanism of the action, since any of the remaining three radicals might be implicated equally. The process can be contrasted sharply with racemization produced by intramolecular rotation.
 - (b) A second example is the pinacone-pinacoline trans-
- ¹ An excellent review of the facts is contained in Porter's " Molecular Rearrangements", New York, 1928.

formation, in which an alkyl radical migrates during a process of dehydration

$$CH_3 \xrightarrow{CH_3} CH_3 \xrightarrow{-H_4O} CH_3 \xrightarrow{CH_3} CH_3$$

$$CH_3 \xrightarrow{CH_3} CH_3$$

$$CH_3 \xrightarrow{CH_3} CH_3$$

$$CH_3 \xrightarrow{CH_3} CH_3$$

This change can be formulated most readily as depending on an *interchange of two radicals on adjacent carbon atoms*, namely, CH₃ and OH, followed by elimination of a molecule of water.

(c) A third example of molecular rearrangement is provided by the *Beckmann change*, in which an oxime is converted into an isomeric amide by acting on it with phosphorus pentachloride and then pouring into water, or by the direct action of a mineral acid

$$\begin{array}{c} C_{\mathfrak{g}}H_{\mathfrak{s}}\cdot C\cdot CH_{\mathfrak{z}} \\ || \\ || \\ || \\ || \\ CH_{\mathfrak{z}}N \end{array} \longrightarrow \begin{array}{c} C_{\mathfrak{g}}H_{\mathfrak{s}}\cdot C\cdot OH \\ || \\ || \\ CH_{\mathfrak{z}}NH \end{array} \longrightarrow \begin{array}{c} C_{\mathfrak{g}}H_{\mathfrak{s}}\cdot CO \\ || \\ || \\ CH_{\mathfrak{z}}\cdot NH \end{array}$$

This change is commonly attributed to an interchange of two radicals on adjacent atoms of carbon and nitrogen, e.g. CH_3 and OH, followed by an $\alpha\gamma$ -migration of a hydrogen atom from oxygen to nitrogen to convert the "enolic" amide into its normal form. The ultimate effect is an isomeric change, but it is possible that, as in the Walden inversion, this depends on substitution and reverse substitution (e.g. of CI for OH and vice versa), accompanied in one case by a change of structure.

The Mechanism of Molecular Rearrangement.

The first two examples of molecular rearrangement are concerned with saturated substances, and even in the Beckmann change the double bond in the oxime appears to play only a secondary part. In each case, therefore, the mechanism is necessarily different from that discussed in the preceding chapter, where the shared electrons of a migrating double bond are drawn upon to neutralize the electric charges of a migrating ion.

In 1924 it was suggested by Lowry 1 that the Walden inversion depends on the momentary formation of a carbonium kation, during an interchange of anions or negatively-charged radicals; and similar views were held as regards the Beckmann change. The widespread acceptance of the view, that molecular rearrangements often depend on the reconstruction of a kation, may be attributed largely to Meerwein, who adopted this view in order to explain his observations on intramolecular displacements of atoms in the camphor series.2 It is, however, important to note that the mere possibility of ionizing an organic chloride or sulphate does not provide in itself any opening for isomeric change, since the reversal of the ionization would only reproduce the original compound. The key to the molecular rearrangement is therefore to be sought and found in the fact that, whereas the separation of a kation from a covalent compound would leave behind a complete octet, the separation of an anion results in the tearing away of a shared duplet, and leaves behind a sextet to act as a centre of instability.3 The molecular rearrangement must therefore be traced to the opportunity of migration afforded by the development of a sextet in the kation, and not merely to the fact that the formation of a chloride or sulphate anion implies the formation of a positively-charged organic kation. The application of this general proposition to individual types of molecular rearrangement is considered in the subsequent sections of the present chapter.

A. The: Walden Inversion

Early Observations.

In 1896, Walden discovered that *l*-malic acid was converted into *d*-chlorosuccinic acid by the action of phosphorus

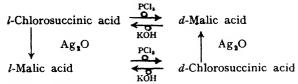
¹ Chemistry and Industry, 1924, 43, 1128.

² Annalen, 1927, 453, 16.

³ Cf. Whitmore, J.A.C.S., 1932, **54**, 3274, 3431, 3435; Whitmore and Stahly, J.A.C.S., 1933, **55**, 4153; Kon, Chemical Society Annual Reports, 1933, **30**, 176.

⁴ Ber., 1896, 29, 133.

pentachloride; but when the product was hydrolysed with silver oxide d-malic acid was regenerated instead of l-malic acid. Regeneration with potassium hydroxide, however, gave back the original l-malic acid. The various interconversions are set out below: ¹



A similar Walden inversion 2 was shown by E. Fischer and Warburg 3 to take place in the conversion of bromopropionic acid into alanine by the action of ammonia and of alanine into bromopropionic acid by the action of nitrosyl bromide.

Br CHMe COOH
$$\rightleftharpoons$$
 NH₂ CHMe COOH Bromopropionic acid.

This inversion can be represented by the following scheme:

$$d$$
-alanine d -Bromopropionic acid d -Alanine.

A Walden inversion, in which the asymmetric carbon atom does not carry a hydrogen atom, was described by McKenzie and Clough in 1910 4 in the case of phenylmethylglycollic acid

¹ The looped arrows indicate the substitutions in which inversion is *now* believed to occur (R. Kuhn and Wagner-Jauregg, *Ber.*, 1928, **61**, 504; Freudenberg and Luchs, *Ber.*, 1928, **61**, 1083).

⁸ E. Fischer, Ber., 1906, 39, 2895.

³ Annalen, 1905, 340, 171. ⁴ J. Chem. Soc., 1910, 97, 1016.

In this case PCl₅ and SOCl₂ gave products of opposite configuration, whereas in Walden's original example of inversion both reagents produced the same reversal of rotatory power in malic acid (and are now believed to produce the same reversal of configuration), although opposite effects were produced by KOH and Ag₂O. The inversion of the corresponding amino-acid was described by McKenzie and Clough in 1912.¹ Walden inversions in a series of secondary alcohols, where the asymmetric carbon atom does not carry either a carbonyl or a carboxyl group, were described by Pickard and Kenyon in 1911 ² and by McKenzie and Clough in 1913,³ e.g.

In this case phosphorus trichloride and hydrogen chloride produced a reversal of sign, but thionyl chloride gave a product of the same sign as the alcohol. On the other hand, silver oxide and soda gave similar products, whereas Walden obtained malic acid of opposite signs by the action of silver oxide and of potash on chlorosuccinic acid.

Direct and Indirect Substitution.

The optical inversion, which results from the successive action of phosphorus pentachloride and of potash on malic acid, shows that a change of configuration must have occurred during the replacement of OH by Cl or conversely. It is, however, a much more difficult task to discover which of the two substitutions which result in a Walden inversion is direct, and which is accompanied by a change of configuration. In particular, this cannot be deduced from a

¹ J. Chem. Soc, 1912 **101**, 390.
² Ibid., 1911, **99**, 45.
³ Ibid. 1913, **103**, 109.

mere inspection of the optical rotations, since substances with similar configurations often have rotatory powers of opposite sign. Thus lævorotatory amyl alcohol yields dextrorotatory esters, although the asymmetric centres are not involved in the process of esterification:

$$\begin{array}{c|c} CH_3 & H & \rightleftharpoons & CH_3 & H \\ C_2H_4 & CH_3 \cdot OH & \rightleftharpoons & CH_3 \cdot O \cdot CO \cdot CH_3 \\ Amyl & alcohol & Amyl & acetate \\ & (dextrorotatory). & (dextrorotatory). \end{array}$$

Many attempts have been made to develop this method of diagnosis by recording the way in which the rotatory power is influenced by changes of wave-length, solvent or concentration, or by the addition of salts, in the hope that similar behaviour in regard to these more complex phenomena would provide evidence of identity of configuration; but all these can be shown to be fallacious if we notice that two compounds of identical configuration, containing the radicals

$$CH_3$$
 and C_3H_7 C

would not only exhibit optical rotations of opposite sign (by reason of the crossing-over from the lighter CH₃ to the heavier C₃H₇ radical), but would be influenced in opposite directions by all the subsidiary changes referred to above.¹

A further complication arises from the fact that the optical inversion is usually incomplete, since a considerable proportion of racemic material is generally produced. This partial racemization is probably due to the simultaneous occurrence of a direct and an indirect action rather than to an independent process of "auto-racemization". Moreover, since we do not know whether a direct or an indirect substitution should be regarded as the more normal, it is not even safe to pronounce in favour of a scheme which reduces the number of inversions to a minimum. This difficulty has been overcome by Kenyon and Phillips, who

¹ Cf. Boys, Proc. Roy. Soc., 1934, A 144, 665, 675.

have been able to detect in which of a pair of substitutions certain Walden inversions occur, by making use of reactions in which the bonds of the asymmetric carbon atom are not involved, so that optical inversion is unlikely to take place. Thus, Phillips in 1923, showed that, when d-methylbenzyl carbinol was converted into the toluene-sulphonic ester by the action of toluene sulphonic-chloride, there was no change in the sign of optical rotation. When, however, the toluene-sulphonic ester was converted into the acetate by the action of potassium acetate, the product was lævorotatory.

(i)
$$d = C_{6}H_{5} \cdot CH_{3}$$
 $CH_{3} \cdot CH_{3} \cdot CH_{3}$
 $C_{6}H_{5} \cdot CH_{3} \cdot CH_{3}$
 $C_{6}H_{5} \cdot CH_{3} \cdot CH_{3}$
 $C_{6}H_{5} \cdot CH_{3} \cdot CH_{3} \cdot CH_{3}$
 $CH_{3} \cdot CH_{3} \cdot CH_{3} \cdot CH_{3}$
 $CH_{3} \cdot CH_{3} \cdot CH_{3} \cdot CH_{3}$
 $C_{6}H_{5} \cdot CH_{3} \cdot CH_{3} \cdot CH_{3} \cdot CH_{3}$
 $C_{6}H_{5} \cdot CH_{3} \cdot CH_{3} \cdot CH_{3} \cdot CH_{3}$
 $C_{6}H_{5} \cdot CH_{3} \cdot CH_{3} \cdot CH_{3} \cdot CH_{3} \cdot CH_{3}$

The acetates, obtained by the substitution (ii), were hydroly ed without change of sign, and therefore gave an alcohol of opposite sign to that from which they were originally prepared.

The esterification (i) does not involve the breaking of the bond between the hydroxylic oxygen and the central asymmetric carbon atom of the alcohol; it was therefore

inferred that the esters must have the same configuration as the alcohols from which they were derived. On the other hand, as Lapworth has pointed out, the hydrolysis of a sulphonate involves the rupture of the link which unites the —O·SO₂·R radical to carbon, whereas the hydrolysis of an acetate depends on the rupture of the bond between RO and CO·CH₂. It can therefore be inferred that optical inversion must have occurred in the replacement of —O·SO₂·C₇H₇ by —O·CO·CH₃, in reaction (ii) and not in the esterification (i) or in the hydrolysis (iii). In this case, then, there could be no doubt as to the stage in which the change of configuration took place. Moreover, no complications arose from the simultaneous occurrence of a direct and indirect substitution, since no indications of racemization were observed.

Mechanism of the Walden Inversion.

Explanations of the Walden inversion which depend on the presence of a hydrogen atom or of a carbonyl-group on the asymmetric carbon atom can be ruled out as not being of general validity, since neither of these radicals is essential to the inversion. Two more general hypotheses, which do not depend on the presence of particular substituents, have been put forward to account for the inversion.

(i) Fischer, Werner, and Pfeisser have supposed that the radicals could push one another about on the surface of a "sphere of affinity", and thus give rise to optical inversion during an interchange of two radicals. Thus, in the reaction

$$C(PQRX) + Y = C(PQRY) + X$$

it was suggested that the place on the surface which was vacated by X might be occupied immediately by Q, so that Y would be obliged to take the place of Q instead of occupying the position vacated by X. According to this hypothesis only one radical in addition to X and Y need

¹ Annalen, 1911, 381, 132.

² Ber., 1911, 44, 881.

³ Annalen, 1911, 383, 123.

be involved in the inversion; and a study of the literature shows that this view has been very widely adopted.

(ii) Gadamer ¹ and Garner ² have suggested that the optical inversion depends on a flattening out of the group C(PQR) on the removal of X, followed by a restoration of the tetrahedral configuration, when Y is added to the other side of this residue. Gadamer ³ represents this optical inversion by the following diagram.

$$d \overset{a}{ \bigodot_{c}^{b}} b \; \leftrightarrows \; \underset{c}{ \hookleftarrow} \; \underset{b}{ \hookleftarrow} \; \overset{a}{ \hookleftarrow} \; \underset{b}{ \hookleftarrow} \; \underset{c}{ \longleftrightarrow} \; \underset{c}$$

On the basis of this general mechanism, he suggested that (i) Attack by a kation (such as the Ag ion of silver oxide) leads to inversion, since in this process the halogen is eliminated from the organic radical, which is thus left free to pick up a hydroxyl ion on the opposite side of the asymmetric carbon atom.

$$\begin{array}{c} a \\ b \\ c \end{array} - C - Cl + \stackrel{+}{A}g \longrightarrow C \stackrel{a}{\smile} b + ClAg.$$

(ii) Attack by an anion (such as the OH ion of an alkali) leads to a direct replacement of Cl by JH, since in this process the Cl atom is progressively pushed aside by the OH ion, without allowing any opportunity for the organic radical to become inverted.

Garner's model represents one of the four bonds of the carbon atom as being drawn through the centre of the atom until it projects out on the other side, whilst the three other

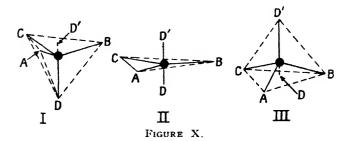
¹ Chem. Zeit., 1912, 36, 1327.

² Proc. Chem. Soc., 1913, 198; cf. Frankland, J. Chem. Soc., 1913, 103, 720.

³ Mitt. pharm. Inst. Univ. Breslau, 1913, 87, 327.

⁴ Loc. cit., pp. 344-5.

bonds are first flattened out, and then reversed, very much like the frame of an umbrella being turned inside out.



"If the model represent a chemical molecule the group D may be replaced by another group with change of configuration. In order to accomplish this according to the Werner theory,1 the entering group must approach the molecule A, B, C, D from the side opposite to that occupied by the group D, and attach itself by means of its partial valencies at D'; if it is attracted to any of the other faces, no inversion would be obtained. Simultaneously with this addition, the group at D is gradually removed, and a corresponding movement of the other valencies occurs in a downward direction. The molecule will pass through an intermediate position II, in which the three valencies A, B and C lie in one plane, and the fourth valency is divided into two halves D and D', and in this position the entering and extruded groups are attached to the carbon atom with equal force. In other intermediate positions, the length of the central valency above and below the ball will indicate the relative strength of the attachment of the two groups. Finally the groups pass into the enantiomorphous position III, in which the group, originally attached at D, is completely removed." 2

It may be said at once that Fischer's explanation is quite incompatible with modern views of valency, since it assumes that the four bonds of a carbon atom, instead of behaving as directed valencies, have the mobility which is characteristic of ions. Moreover, if the bonds were really mobile, racemization would be more likely to occur than optical inversion. Gadamer's hypothesis also implies a breaking

¹ Ber., 1911, 44, 881.

² Garner, Proc. Chem. Soc., 1913, 198-201.

away from the stereochemical conception of van't Hoff, for which at the time no justification whatever could be given. The experimental work of the last twenty years has proved, however, that the carbon atom has not one but two stable configurations, namely (i) a tetrahedral configuration, which occurs in diamond and in all quadricovalent carbon compounds, and (ii) a triangular configuration, which occurs in graphite and in tercovalent carbonium (kat)ions. Thus X-ray analysis has not only confirmed the tetrahedral configuration of the carbon atoms in diamond and in the zigzag chains of hydrocarbons and the hexagonal structure of the rings in graphite, but it has shown that the three oxygen atoms in the carbonate ion of calcite and aragonite are distributed symmetrically around the central atom of carbon in a planar or triangular configuration.

$$\vec{O}$$
 compare \vec{O} \vec{O}

If, therefore, we accept the view that the Walden inversion depends on an *interchange of anions*, the mechanism suggested by Gadamer and the "umbrella" model of Garner can be rationalized, as Lowry suggested in 1924, in such a way as to involve no *ad hoc* hypotheses at all, since ample precedents can be produced for every stage that need be postulated.

This modified mechanism can be illustrated very readily in the inversions studied by Phillips and his colleagues, namely

² J. Chem. Soc., 1923, 123, 44.

¹ Chemistry and Industry, 1924, 43, 1128; cf. Second Solvay Report, 1925, 40.

$$(ii) C_{\mathbf{s}}H_{\mathbf{5}}O \cdot CO \cdot CH_{\mathbf{3}} + \dot{K}O \cdot CO \cdot CH_{\mathbf{3}} \cdot C_{\mathbf{7}}H_{\mathbf{7}} + \dot{K}O \cdot CO \cdot CH_{\mathbf{3}} \cdot C_{\mathbf{7}}H_{\mathbf{7}}.$$

In these inversions we must postulate (i) that the sulphonate radical separates from the ester in the form of an anion, since this is the form in which it appears in the final products of the reaction, (ii) that the residue left behind by the removal of the sulphonate ion must therefore be positively-charged, i.e. it consists of a tercovalent carbonium ion, which either has or very readily assumes a planar configuration, (iii) that the tercovalent carbonium ion finally unites with an acetate ion to give an ester, in which the central atom of carbon is again quadricovalent, and therefore once more assumes a tetrahedral configuration.

The essential stage in the inversion is that in which a sulphonate ion separates from the radical, e.g. on the lefthand side, whilst an acetate ion unites with the radical on the right-hand side, thus

No special hypothesis is then needed, since (i) the conversion of a covalence into an electrovalence, and conversely, has already been postulated in many other cases, and (ii) the planar character of the tercovalent ion has already been postulated and proved in the calcite ion of calcite and aragonite.

The mechanism suggested above indicates that (i) Direct substitution, without change of configuration, will occur

whenever one anion replaces another on the same side of the kation, as in the scheme

Direct substitution would also occur in any process which did not involve the conversion of the asymmetric atom into a positively-charged ion. (ii) Racemization would occur inevitably if the planar carbonium ion should become completely free, since it would necessarily yield equal quantities of the two optical antipodes on resuming the tetrahedral configuration. Partial racemization would also result if the direct substitution and the indirect substitution proceeded as simultaneous reactions. (iii) Walden inversions will follow whenever the two anions which are interchanged lie on opposite sides of the planar kation at the moment of interaction.

It is not possible to predict from this mechanism which reactions will give rise to normal substitution and which to a Walden inversion; but certain factors which may favour or inhibit inversion are indicated below.

- (a) Steric Hindrance. In discussing this mechanism it was suggested by Lowry "that two bulky multipolar anions might find it easier to approach a planar kation on opposite sides than on the same side of the plane of symmetry". This steric factor must certainly play a part in controlling the course of the substitution and is perhaps adequate to account for the quantitative inversion which accompanies the displacement of a sulphonate by an acetate radical.
- (b) Formation of Polarized Intermediate Compounds. Kenyon and his colleagues have described a number of reactions in which the occurrence of a Walden inversion can be associated with the formation of an intermediate complex carrying a positive charge on the substituent, together with liberation of a chloride anion, which

¹ J. Chem. Soc., 1930, 415; 1931, 382.

is thus given an opportunity to approach the organic radical on the opposite side. The following examples are given:

(i) The interaction of l- β -octanol with hydrogen chloride to form d- β -chloro-octane ¹

$$\begin{array}{c|c}
 & C_{\bullet}H_{13} \\
 & CH_{3}
\end{array}$$

(ii) The interaction of l- β -octyl chlorocarbonate and pyridine ²

$$\begin{array}{c|c} L & C_{\mathfrak{g}}H_{13} & H \\ \hline CH_{\mathfrak{g}} & C_{\mathfrak{g}}CO\cdot CC \\ \hline & CH_{\mathfrak{g}} & C_{\mathfrak{g}}H_{13} \\ \hline & CH_{\mathfrak{g}} & C_{\mathfrak{g}}H_{13} \\ \hline & CH_{\mathfrak{g}} & C_{\mathfrak{g}}H_{13} \\ \hline & CH_{\mathfrak{g}} & CH_{\mathfrak{g}} \\ \hline & CH$$

(iii) The interaction of ethyl *l*-mandelate with phosphorus pentachloride

$$\begin{array}{c|c}
C_{e}H_{b} & C_{e}H_{b} \\
\hline
EtO\cdot CO & C_{e}H_{b} \\
\hline
C_{e}H_{b} & C_{e}H_{b}$$

(iv) The interaction of ethyl *l*-mandelate with thionyl chloride in presence of pyridine

In the absence of pyridine a normal substitution occurs.

¹ J. Chem. Soc., 1911, 99, 45, cf. ibid., 1929, 1700.

² Ibid., 1929, 2510.

338 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

(v) The conversion of l- β -octanol into d- β -chloro-octane by thionyl chloride both in the absence 1 and in the presence 2 of pyridine. The initial interaction does not touch the asymmetric atom, and is therefore unlikely to be accompanied by inversion; it is therefore suggested that this may occur when the intermediate chlorosulphinate is decomposed:

(c) Sign of Ionization of the Organic Radical.

Gadamer suggested that a kation, such as Ag, gave rise to inversion, whilst an anion, such as OH, gave rise to direct substitution. This preliminary suggestion has not been confirmed, since it is now believed that in Walden's classical example, inversion occurs during the action of potash but not when silver oxide is used, although in the aromatic series both reagents produce the same effect in the reaction

Kenyon, Lipscomb and Phillips 4 have suggested, however, that certain aromatic compounds, instead of breaking down into an organic kation and an inorganic anion, and so permitting inversion to take place, may break down into an inorganic kation and an organic anion, which would carry a complete octet of valency electrons and would thus be maintained in a tetrahedral configuration, in which no inversion could take place. Thus it is suggested that the

¹ Levene and Mikeska, J. Biol. Chem., 1924, 59, 45.

² McKenzie and Tudhope, ibid., 1924, 62, 551.

⁸ McKenzie, J. Chem. Soc., 1913, 103, 687.

⁴ J. Chem. Soc., 1930, 422.

elimination of sulphur dioxide from an intermediate chlorosulphinate may proceed as follows:

Reaction B is rendered possible by the electron-attracting properties of the phenyl-group, whereas purely aliphatic groups are electron-donating. In presence of pyridine a pyridinium complex is formed, which carries a positive charge, and so attracts electrons away from the asymmetric carbon atom and prevents it from forming an anion

A Walden inversion therefore occurs, which does not take place in the absence of pyridine, even when potassium carbonate is added. In the case of d-\beta-octanol, where there is no aromatic nucleus, an inversion accompanies the action of thionyl chloride both in the absence and in the presence of pyridine.

Evidence that carbonium anions are actually stable may be cited from two sources.

(i) In 1927, Kuhn and Albrecht 1 discovered that secondary nitrobutane retains its optical activity after conversion to the sodium salt.

$$CH_s$$
 C_sH_s
 C_sH_s
 C_sH_s
 C_sH_s
 C_sH_s
 C_sH_s
 C_sH_s

This observation has been confirmed by Shriner and Young,² and can only be explained on the assumption that the tercovalent carbon anion which also carries a "lone pair"

¹ Ber., 1927, **60**, 1297.
² J.A.C.S., 1930, **52**, 3332.

of electrons, retains its tetrahedral configuration 1 and therefore its optical activity.

(ii) E. S. Wallis and F. H. Adams in 1933 have shown that an optically-active carbon anion, in the compound (I) can exist for an appreciable length of time.

On the other hand, the optically-active covalent compound (II) was immediately racemized on conversion to (III) in which the tercovalent carbon is positively charged

It thus appears that carbon compounds, like those of nitrogen and sulphur, retain their spatial configurations as lone as each atom still possesses a complete octet of electrons, but that the tetrahedral configuration becomes unstable in tercovalent carbonium compounds in which the carbon atom possesses only six valency electrons. The Walden inversion must therefore be limited to an exchange of anions, which converts the asymmetric carbon atom (at least momentarily) into the form of a kation.

(d) **Dipole Moments.** An examination of the dipole moment, and its vector orientation, in molecules which

¹ Cf. Mills, British Association Report, 1932; see also Nature, 1932, 130, 566.

² J.A.C.S., 1933, **55**, 3838.

can undergo the Walden inversion would undoubtedly yield results of direct significance, since the direction of approach of a reactant anionoid group is undoubtedly conditioned by the vector disposition of the local electrical fields which surround the asymmetric atom.¹ For instance, the different behaviour of aliphatic and aromatic α-hydroxy acids may be connected with the difference of vector sign of the moments in the groups CH₃—C and Ph—C (cf. p. 68).

Racemization of Compounds in Which Free Rotation is Spatially Restricted.

The Walden inversion, which can only occur during a definite chemical reaction, can be sharply contrasted with internal racemization produced by the intramolecular rotation of radicals in an initially dissymmetric molecule.

It has been pointed out that Werner's concept of free interchange in position of radicals attached to a carbon atom is not in accordance with modern ideas of covalency. However, it has been a fundamental assumption of organic chemistry that two atoms joined by a single covalent bond can have *free rotation* about that bond as an axis. This has been supported by the theoretical calculations of Pauling.²

Bischoff 3 in 1891 suggested that complete rotation of carbon atoms could be restricted by the presence of substituents on them, but not until 1922 was any experimental evidence brought forward to support that view. In that year Christie and Kenner 4 resolved 6:6'-dinitrodiphenic acid (I) into optically active enantiomorphs, and found that the active acids were remarkably stable. They suggested that the active diphenic acids possessed coaxial but

¹ Cf. Polanyi, Proc. Roy. Soc., 1934, **B, 116,** 202: Mem. Manchester Phil. Soc., 1934, **41**, 41. Ogg and Polanyi, Trans. Faraday Soc., 1935, **31**, 604. Hughes, Juliusburger, Topley and Weiss, J. Chem. Soc., 1935, 1525, 1936, 1173.

² J.A.C.S., 1931, **53**, 1367, 3229.

³ Ber., 1891, 24, 1085. ⁴ J. Chem. Soc., 1922, 121, 614.

342 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

not coplanar rings, free rotation of which about the linking bond was restricted by the mutual interaction of the substituent groups.

Almost simultaneously in 1926, Le Fèvre and Turner, Bell and Kenyon, and Mills, pointed out that the substituent groups in the *ortho*-position to the linking bond of any diphenyl derivatives would, if large, spatially obstruct one another, and thus produce molecular dissymmetry. Other reactions (e.g. polar attractions or repulsions) between the substituent groups might, however, produce, or enhance, the same resultant effect.

It was pointed out that in general three ortho-substituent groups would be needed to produce steric blocking, but that, if sufficiently large, two ortho-substituent groups might suffice. This has been confirmed experimentally. Thus diphenic acid is not resolvable, but 6-nitrodiphenic acid can be resolved. 2: 2'-Di-iodo-4: 4'-dicarboxyldiphenyl (II), b' diphenyl-2: 2'-disulphonic acid (III) and the methiodide of 2: 2'-bisdimethylamino-diphenyl (IV) have all been resolved, showing that two groups, if sufficiently large, can prevent free rotation in the diphenyl system. Similar examples of dissymmetry due to restricted rotation about a single bond have been proved to exist in the phenyl-pyrrole and dipyrryl systems.

¹ Chemistry and Industry, 1926, 45, 831.

² Ibid., 1926, **45**, 864. ³ Ibid., 1926, **45**, 884, 905.

⁴ Bell and Kenyon, Chemistry and Industry, 1926, 45, 864.

⁵ Searle and Adams, J.A.C.S., 1933, 55, 1649.

⁶ Lesslie and Turner, J. Chem. Soc., 1932, 2394.

⁷ Turner, Chemistry and Industry, 1932, 51, 435.

⁸ Bock and Adams, J.A.C.S., 1931, 53, 374.

Note.—A detailed review of experimental work is given by R. Adams and Yuan, Chemical Reviews, 1933, 12, 261-338.

COOH

I

SO₃H

$$N(CH_3)_3I$$

COOH

 $COOH$
 $IIII$
 $IIII$
 $IIII$
 $IIII$
 IIV

In nc instance has it been possible to connect other factors than spatial dimensions with the inhibition of free rotation. Further, X-ray measurements of molecular dimensions, and measurements of dipole moments both support the stereochemical explanations for these cases of molecular dissymmetry. From X-ray data it is easily possible to calculate the minimum dimensions of ortho-substituent groups which could prevent free rotation in the diphenyl system and from such calculations 1 one can predict which substituents will make a diphenyl derivative resolvable. experimental results are fully in accord with the theoretical calculations, but it has been found that those diphenvl derivatives in which the dimensions of the ortho-substituents are but slightly greater than the minimum for the prevention of free rotation are all substances which racemize with ease. Thus whereas active 6: 6'-dinitrodiphenic acid is very stable, 6-nitrodiphenic acid may be racemized on prolonged boiling of solutions of its salts and 2: 2'-diffuoro-*6: 6'-diaminodiphenyl (V) racemizes rapidly.

$$\underbrace{ \bigvee_{NH_{\underline{s}}}^{F} \bigvee_{NH_{\underline{s}}}^{F}}_{NH_{\underline{s}}}$$

Kuhn and Albrecht² have shown that the decrease of optical activity of diphenyl derivatives, which is accelerated

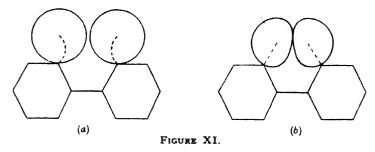
¹ Cf. Stanley and Adams, J.A.C.S., 1930, **52**, 1200.

² Annalen, 1927, 455, 272; 458, 221; 1928, 464, 91; 465, 282; Kuhn, 4th Solvay Report, 1931.

344 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

by heating, follows a unimolecular law. It seems, therefore, that free rotation can be effected in those molecules of the active substance which possess more than a definite minimum energy content.¹ One may suppose either that rapidly rotating molecules can be deformed sufficiently for the steric interaction to be overcome, or, alternatively, that the molecules momentarily dissociate into portions which reunite after brief independent movement has occurred.

It may be pointed out that valency deflection alone can produce sufficient deformation of molecular structure to effect the racemization of an active diphenyl derivative (see Figure XI (a)), and that the variability of valency angles has been proved by numerous measurements of the halogen derivatives of methane, benzene and naphthalene.



Kuhn has pointed out 2 that the formation of fluorenone carboxylic acid (VI) by heating diphenic acid in concentrated sulphuric acid solution proves that valency bonds can be bent.

² 4th Solvay Report, 1931.

¹ Cf. Sidgwick, Chemical Society Annual Reports, 1931, 28, 70.

Further, all molecules are compressible, and the local compression of two radicals on impact could effect race-mization, provided that the mutual kinetic energy was sufficient. (Figure XI (b)).

It is therefore highly probable that in the case of the racemization of derivatives of diphenic acid one is not dealing with a chemical reaction similar to that involved in the formation of racemic acid when d-tartaric acid is boiled with alkali.

The active diphenic acids, of suitable structure, can be racemized by heating in neutral, or indifferent solvents, though with 6-nitrodiphenic acid mutarotation occurs more quickly in a glass than in a quartz polarimeter tube, indicating either that alkali is a catalyst or else that the anion may be more mobile than the free acid.

By comparison of the relative speeds of racemization of different diphenyl derivatives it has been possible to obtain an estimate of the relative dimensions of substituent groups, which is consistent throughout with the computations from X-ray data. It appears that the relative order of dimensions of substituent groups is

hydrogen fluorine hydroxyl methoxyl amino carboxyl.

For derivatives of 2-nitro-6-carboxydiphenyl the following figures were obtained.

TABLE I

Other substituents.		Solven	Half life period.		
2'-Br, 5'-CH ₈	Acetic	acid	at	118°	3,240 min.
2'-CH ₃	,,	,,	,,	,,	179 ,,
2'-Cl, 5'-CH ₃	,,	,,	,,	,,	154 ,,
2'-NO,	,,	,,	,,	,,	125 ,,
2'-COOH	,,	,,	,,	,,	91 ,,
2'-OCH,	Abs. I	EtOH	at	25°	9.4 ,,

The effect of size in preventing free rotation is therefore predominant. Polar effects, however, may not be completely absent, for a substituent in the meta-position to the

¹ Kuhn and Wagner-Jauregg, Naturwissenschaften, 1929, 17, 103.

Adams and Yuan, Chemical Reviews, 1933, 12, 261-338.

bond about which rotation is restricted can affect the ease of racemization.¹ This meta-substituent may act sterically by altering valency angles throughout the whole molecule and consequently the inter-atomic distances, or again it might affect the rotational energy content of particular parts of the molecule. The precise action is still obscure, but it has been pointed out that atomic weight, and consequently rotational momentum, is not a factor entering into consideration, since substituent chlorine and bromine atoms have precisely the same effect in slowing down the racemization of diphenyl derivatives.

Table II

Half life periods at 26° C. of the active forms of substituted 2-nitro-6-carboxy-2'-methoxydiphenyls and their salts

				Half life Period in Minutes.		
Compound.				Free Acid in Alcohol.	Brucine Salt in CHCl ₃ .	
No substit	uent			9.4	12.0	
5'-OCH ₃				10.8	16.7	
5'-CH ₃				11.5	15.8	
5'-Cl .		•		31.0	16.7	
5'-Br .		•		32.0	16.7	
5'-NO ₂				35.4	7.7	

The suggestion that valency deflection must be taken into account in explaining the racemization of these active diphenic acids is supported by the fact that although stereochemical models indicate that many derivatives of diphenylmethane, benzophenone, diphenylsulphone and similar substances should exhibit the same effects of steric interference yet attempts to prepare optically active derivatives from any of these ring structures have been unsuccessful.

B. THE PINACONE—PINACOLIN AND RELATED TRANSFORMATIONS 2

(i) The Pinacone—Pinacolin Transformation.

When pinacone (I) is treated with a strong mineral acid it is converted into pinacolin (II) by (a) loss of a molecule

¹ Adams and Yuan, J.A.C.S., 1932, 54, 2966, 4434.

² A detailed account is given in J. W. Baker's "Tautomerism", (Routledge, London, 1934), Chap. XIV.

of water and (b) migration of a methyl group from one carbon atom to an adjacent one.

$$(I) \xrightarrow[CH_3]{CH_3} \xrightarrow[CH_3] \xrightarrow[CH_3]{CH_3} \xrightarrow[CH_3]{CH_3} \xrightarrow[CH_3]{CH_3} \xrightarrow[CH_3]{CH_3} \xrightarrow[CH_3] \xrightarrow[CH_3]{CH_3} \xrightarrow[CH_3]{CH_3} \xrightarrow[CH_3]{CH_3} \xrightarrow[CH_3]{CH_3} \xrightarrow[CH_3]$$

The change is a general one for all compounds of this type, and, in the case of unsymmetrical pinacones, it is usually found that one hydrocarbon group is more mobile than others. Tiffeneau ¹ has, in a long series of researches, investigated this order of mobility with the object of ascertaining, through it, the relative stabilities of different hydrocarbon radicals. After detailed study of the velocity of the reaction under different conditions, he came to the conclusion ² that the initial chemical change was the removal of a molecule of water from that of the pinacone, and represented the resulting product as a diradical possessing two free valencies, which could only be satisfied by the migration of a hydrocarbon group.

When it is recalled that the molecular rearrangement is effected by the use of a strong mineral acid it is more reasonable to suppose that the reaction is initiated by the removal of a hydroxyl ion. Thus, sulphuric acid by promoting the dissociation of hydroxyl ion from the initial glycol (I) would strengthen the opposing electrical fields due to the dipoles

in the groups C-O-H, which are so disposed that adjacent carbon atoms are centres of positive polarization.

The migration of a methyl group as an anion will relieve the electrical instability of (Ib), and, by the final elimination of a proton from the second hydroxyl group, the molecule can then acquire an electrically stable configuration.

¹ Tiffeneau, 2nd Solvay Reports, 1925, pp. 291-7.

² Compt. rend., 1906, 143, 687.

348 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

This scheme, suggested by Ingold and Shoppee, can be applied equally well to many other cases of migration of hydrocarbon radicals. Some typical examples are mentioned in the following sections.

(ii) Semipinacolinic Transformations.

The pinacone change is not confined to tertiary glycols. Thus Tiffeneau a has carried out the following transformations:

McKenzie has shown that a similar migration occurs when an $\alpha\beta$ -amino-alcohol is treated with nitrous acid.³ In the case of the glycol, one of two alternative hydroxyl groups may dissociate as hydroxyl ion, and in consequence here results, in general, a mixed product in which either the aldehyde (II) or the ketone (III) preponderates according to the nature of the transforming reagent.

In the reaction of nitrous acid with the amino-alcohol (IV), however, only the ion (V) can be produced, and hence a hydrocarbon group invariably migrates to the carbon atom from which the amino group has been detached.

¹ J. Chem. Soc., 1928, 371: cf. Ingold, Annual Reports, 1928, p. 133; see also Robinson, Mem. Manchester Phil. Soc., 1920, 64, iv, 7; Kermack and Robinson, J. Chem. Soc., 1922, 121, 440.

² Bull. Soc. Chim., 1923, (iv), 33, 735.

³ J. Chem. Soc., 1924, 125, 2105.

$$CH_{3} \xrightarrow{C_{6}H_{5}} H$$

$$CH_{3} \xrightarrow{C_{6}H_{5}} CC_{6}H_{5} \xrightarrow{C_{6}H_{5}} CC_{6}H_{5}$$

$$OH \xrightarrow{NH_{2} + O = N - O - H} OH \xrightarrow{+ N_{2} + OH^{-}} (v) \xrightarrow{+ H_{2}O}$$

$$CH_{3} \xrightarrow{C_{6}H_{5}} CC_{6}H_{5}$$

$$CH_{3} \xrightarrow{C_{6}H_{5}} CC_{6}H_{5}$$

(iii) The Benzil--Benzilic Acid Transformation.

Benzil (VII) is, like pinacone, a molecule in which the internal dipole moments are opposed to one another. By the action of warm caustic alkali it yields, with addition of the elements of water, benzilic acid (VIII), in which the internal electrostatic strain has been relieved.

This change can best be explained by postulating the migration of a negative phenyl radical from one carbon atom to the adjacent one. Ingold 1 has suggested that the change is initiated by the addition of hydroxyl ion to the carbon atom of one of the carbonyl groups:

whilst, conversely, Whitmore 2 assumes that the initial

¹ Chemical Society Annual Reports, 1928, 25, 124.

² J.A.C.S., 1932, **54**, 3274; 1933, **55**, 4153.

stage is the addition of a positive ion to the oxygen of a carbonyl group:

Whitmore's representation indicates that the reaction could be brought about by the augmentation of an internal polar field, but on the other hand it does not explain why the change requires the presence of alkali.¹

(iv) The Wagner-Meerwein Rearrangement.2

Many molecular rearrangements analogous to those described above have been discovered in the course of investigations of naturally-occurring terpenes. Thus, on treatment with hydrogen chloride, pinene (IX) is converted into bornyl chloride (X), though pinene hydrochloride (XI) can be obtained if the reaction is carried out in a freezing mixture.

¹ Cf. Chemical Society Annual Reports, 1933, 30, p. 177.

² Cf. J. W. Baker, "Natural Terpenes", 1930, pp. 89-110; Tautomerism", 1934, Chap. XIV.

Similarly, esters of camphene hydrate, on warming in solution, give *iso*bornyl esters.

Meerwein 1 has demonstrated that these reactions are essentially rearrangements of hydrocarbon kations, similar in type to the pinacone—pinacolin change. The molecular rearrangement of the kation of the terpene will undoubtedly occur in such a manner as to stabilize the molecule. Both the disposition of the small local electrical dipoles within the molecules and the spatial proximity of groups must be taken into account if one is to attempt to explain why a particular change occurs. It is of interest to note that very frequently a tertiary ester (e.g. pinene hydrochloride) isomerizes to a secondary ester (e.g. bornyl chloride), and that the elimination of a molecule of water or of hydrogen chloride results in the formation of an ethylenic linkage rather than a strained homocyclic ring.

C. THE BECKMANN TRANSFORMATION

Isomeric Change of Oximes.

In 1886 Beckmann ² discovered that, when oximes were treated with acetic anhydride or acetyl chloride, they were not always acetylated, but often underwent isomeric change and were thus converted into substituted amides. He showed subsequently that many reagents of acidic character, such as concentrated sulphuric acid and phosphorus pentachloride, could bring about the same isomeric change. Additional interest arose from the observation that the α - and β - forms of certain oximes yielded different amides.

$$\alpha \xrightarrow{C_6 H_5} C = \text{NOH} \longrightarrow C_6 H_5 \cdot \text{NH} \cdot \text{CO} \cdot \text{CH}_3$$

$$\beta \xrightarrow{C_6 H_5} C = \text{NOH} \longrightarrow C_6 H_5 \cdot \text{CO} \cdot \text{NH} \cdot \text{CH}_3$$

¹ Ber., 1921, 53, 1815; Annalen, 1927, 453, 16.

² Ber., 1886, 19, 988; 1887, 20, 1507, 2580; 1888, 21, 516.

At that time the reason for the occurrence of isomeric oximes was unknown. In 1891, however, Hantzsch 1 suggested that the three valencies of the nitrogen atom were not necessarily coplanar, but might be arranged like three of the four bonds of a carbon atom, and could thus give rise to a stereoisomerism comparable with that of fumaric and malic acid. He suggested also that the configuration of ketoximes could be deduced from the structure of the amides formed from them in the Beckmann rearrangement. For this purpose he assumed that, under the influence of the reagent, vicinal groups changed places, giving a labile form of the amide, which then reverted to the stable form, thus:

In the same way, it was suggested that the configuration of the aldoximes could be deduced from the fact that one of them was acetylated by the action of acetic anhydride, whilst the other (which he therefore formulated as the ciscompound) was dehydrated to the nitrile,

Interchange and Elimination of Groups in trans-Positions.

Subsequent researches have shown that the method of diagnosis suggested by Hantzsch is not only arbitrary but erroneous, since trans-elimination of groups from a molecule is actually more frequent than cis-elimination; in particular, the interchange and elimination of groups in the oximes takes place quite regularly in the trans-positions. Thus Meisenheimer 2 in 1921 showed that ozonolysis of triphenylisooxazole at o° gave the benzoyl derivative of

¹ Ber., 1891, 24, 22. ² Ibid., 1921, 54, 3195, 3208.

 β -benzil monoxime, in which oxime a phenyl group migrates during the Beckmann rearrangement:

This conclusion was later substantiated by effecting the following changes ¹

The trans-elimination of groups in the aldoximes was established in a similar way by Brady and Bishop,² thus:

¹ Annalen, 1926, **446**, 205. ² J. Chem. Soc., 1925, **127**, 1357.

The β -oxime loses water and gives a nitrile, but the halogen is also eliminated, perhaps through the formation of an unstable isooxazole. The α -oxime, in which the hydroxyl group is directed away from the ring, yields an acetate from which it can be regenerated, in spite of the fact that the hydrogen and hydroxyl radicals are favourably placed for cis-elimination.

Mechanism of the Beckmann Transformation.

The reagents that bring about the Beckmann transformation are all of an acidic nature, i.e. they are reagents with the help of which hydroxyl can be replaced by a more negative radical. Thus in the changes first studied by Beckmann the hydroxyl group was acted on by acetylating agents; and the subsequent use of sulphuric acid and phosphorus pentachloride can be explained as depending on the conversion of the oxime into a sulphate or chloride. It is therefore natural to adopt the view that the transformation depends on ionization, and picture the change as involving three essential steps: (i) Ionization, resulting from the replacement of the hydroxyl radical by a negative radical. This radical has a greater electron-affinity than hydroxyl, and this enables it to seize the shared electrons of the bond and thus to form a negatively-charged ion.

$$\begin{array}{c} \text{R:N-OH} \xrightarrow{\text{Ac}_1\text{O}} \text{R:N-O-CO-CH}_3 \rightleftharpoons \text{R:N} \xrightarrow{\text{O}} \text{CO-CH}_3 \\ \\ \text{R:N-OH} \xrightarrow{\text{H}_1\text{SO}_4} \text{R:N-O-SO}_3 \text{-OH} & \rightleftharpoons \text{R:N} \xrightarrow{\text{O}} \text{O-SO}_3 \text{-OH} \\ \\ \text{R:N-OH} \xrightarrow{\text{PCI}_4} \text{R:N-Cl} & \rightleftharpoons \text{R:N} \xrightarrow{\text{C}} \text{Cl} \end{array}$$

¹ Cf. Kuhara and Todo, Mem. Coll. Sci. Eng. Kyoto, 1910, 2, 387; Stieglitz in Porter's "Molecular Rearrangements" (New York, 1928), p. 42, reverts to a system in which "plus and minus signs no longer imply a complete transfer of an electron from one atom to another" but "indicate merely a displacement of the pair of valence electrons from the central position". The formulæ which are used to express this point of view do not convey any clear impression of the polarity of the molecule, since the multivalent atoms usually carry both positive and negative signs simultaneously.

(ii) Rearrangement of the kation, including in particular the migration of an alkyl or aryl radical to heal the scar left by the separation of the anion. This arrangement can be attributed to the instability of the positive radical R=N, in which the nitrogen carries only a sextet of electrons. In order to be effective, the alkyl or aryl radical must migrate as an anion, with a complete octet of electrons. It then carries a negative charge which serves to neutralize the positive charge on the nitrogen, e.g.,

$$\begin{array}{ccc}
C_{\bullet}H_{\bullet}\cdot C\cdot CH_{3} & \xrightarrow{C_{\bullet}H_{\bullet}\cdot C} \\
\parallel & & \parallel \\
N + & & N\cdot CH_{3}
\end{array}$$

It must be admitted that the rupture of a carbon-to-carbon bond is much less easy than the dissociation of a proton in the prototropic changes of the preceding chapter; but there is abundant evidence to show that migrations of this kind can occur under the more drastic conditions now under consideration. In this connection attention may be directed to the migration of a phenyl group when triphenylmethyl peroxide is converted by heating into a pinacone.¹

(iii) Elimination of the Anion. The migration of a negatively-charged alkyl or aryl radical leaves behind a positive charge on the carbon atom of the oxime group, which is left with only a sextet of electrons. Carbon is, however, less ready than nitrogen to form salts and is not easily retained in the form of a carbonium ion. When, therefore, in the last stage of the transformation, the product is brought into contact with water, the anion is finally eliminated and replaced by hydroxyl, forming a covalent compound of quadrivalent carbon, e.g.,

$$\begin{array}{ccc} C_0H_5\cdot \overset{+}{C} & \overset{-}{C}I + \overset{+}{H_2O} & \xrightarrow{C_0H_5\cdot C\cdot OH} + \overset{+}{HCI} \\ \parallel & & \parallel & & \\ N\cdot CH_3 & & & NCH_3 \end{array}$$

¹ Wieland, Ber., 1911, 44, 2550.

(iv) Migration of a Proton. The final stage is a migration of a proton to give an amide in its ordinary stable form. In this process the newly-formed bond between carbon and oxygen persists, but the link between oxygen and hydrogen is ionized by the separation of a proton.

Since this stabilization is only possible when the negative radical has been replaced by hydroxyl, it may be a factor in promoting the final elimination of the anions shown under (i) above.

The Beckmann Rearrangement as an Isomeric Change.

The mechanism suggested above has been confirmed by the experiments of Chapman and Howis ¹ on the spontaneous rearrangement of the picryl ethers of certain oximes. Chapman took advantage of the fact that the negative character of the picrate radical is insufficient to bring about a spontaneous isomeric change during the preparation of the picrate of the oxime by a modified Schotten-Bäumann method, but is nevertheless sufficient to produce a spontaneous rearrangement on heating. Thus benzophenone-oxime picryl ether (I) melts almost explosively when heated to 106°, but re-solidifies as benz-N-picryl-nilide (III), which melts again at about 197°.

$$\begin{array}{c} C_{\mathfrak{g}}H_{\mathfrak{s}}\cdot C \cdot C_{\mathfrak{g}}H_{\mathfrak{s}} \\ \parallel \\ N \cdot OC_{\mathfrak{g}}H_{\mathfrak{s}}(NO_{\mathfrak{s}})_{\mathfrak{s}} \\ \parallel \\ C_{\mathfrak{g}}H_{\mathfrak{s}}\cdot N \end{array} \begin{array}{c} C_{\mathfrak{g}}H_{\mathfrak{s}}\cdot C \cdot O \cdot C_{\mathfrak{g}}H_{\mathfrak{s}}(NO_{\mathfrak{s}})_{\mathfrak{s}} \\ \parallel \\ C_{\mathfrak{g}}H_{\mathfrak{s}}\cdot N \end{array} \begin{array}{c} C_{\mathfrak{g}}H_{\mathfrak{s}}\cdot C = O \\ C_{\mathfrak{g}}H_{\mathfrak{s}}\cdot N \cdot C_{\mathfrak{g}}H_{\mathfrak{s}}(NO_{\mathfrak{s}})_{\mathfrak{s}} \\ C_{\mathfrak{g}}H_{\mathfrak{s}}\cdot N \cdot C_{\mathfrak{g}}H_{\mathfrak{s}}(NO_{\mathfrak{s}})_{\mathfrak{s}} \\ \parallel \\ C_{\mathfrak{g}}H_{\mathfrak{s}}\cdot N \cdot C_{\mathfrak{g}}H_{\mathfrak{s}}(NO_{\mathfrak{s}})_{\mathfrak{s}} \\ \parallel \\ C_{\mathfrak{g}}H_{\mathfrak{s}}\cdot N \cdot C_{\mathfrak{g}}H_{\mathfrak{s}}(NO_{\mathfrak{s}})_{\mathfrak{s}} \\ \parallel \\ N \cdot Phenylbenziminopicryl ether. \\ \parallel \\ m.p. \ 107^{\circ}). \end{array}$$

The intermediate product (II) is too unstable to be isolated; the velocity of the isomeric change is therefore controlled by the initial Beckmann transformation, I—II, which proceeds as a unimolecular reaction, in which half the material

is changed in about 80 minutes in chloroform at 50°. The velocity constants in a series of solvents at 50° were as follows:

$$C_0H_0$$
 CHCl₃ $C_2H_4Cl_2$ $k \times 10^8 = 0.24$ 4.0 8.35

Similar measurements of velocity were made with the picryl ethers of acetophenone-oxime

$$C_{\bullet}H_{\bullet}$$
 $C=N\cdot O\cdot C_{\bullet}H_{\bullet}(NO_{\bullet})_{\bullet};$

and measurements at temperatures of 100° and 116° showed that the energies of activation for the isomeric change of the two ethers were about 27,000 cal. per gram molecule. It was also shown that the picrates of the stereo-isomeric α - and β -oximes of p-chlorobenzophenone underwent the same Beckmann change as the oximes themselves when acted on successively by phosphorus pentachloride and water.

The rearrangements did not appear to require an external catalyst, since the velocities were not affected by repeated purification of the ethers or of the solvents, or by the addition of broken glass. Moreover, the velocities were not affected by carrying out the transformations in a blackened tube, and no rearrangement was brought about by irradiation with a mercury lamp.

From velocity measurements with benzophenone-oxime picryl ether in carbon tetrachloride solution 1 it was discovered that the concentration of the solution had an important influence on the rate of the rearrangement. The velocity constant (k) at any one temperature can be expressed by the equation $k = k_0 + xc$, where k_0 is the velocity coefficient at infinite dilution and x the increase of velocity for a concentration of one gram molecule per litre. The velocity rate indicated that, at infinite dilution, the oxime ether molecules could be activated by collision

with solvent molecules, but that, in concentrated solutions, activation could also be brought about by collision between two molecules of ether, or between a molecule of ether and a molecule of the rearrangement product. The activation energy for the solvent collision is 30,250 cal. and that for the collision between two molecules of ether 22,620 cal. per gram molecule. The difference in activation energy of the two processes is considerable, and hence it must be concluded that a molecule of oxime ether can be more easily activated when in the vicinity of a molecule of the same type thar when it is surrounded completely by inert carbon tetrachloride molecules. This fact is easily explicable, since the dissociation of the oxime ether involves the development of electrical charges at the ends of the bonds concerned, and less energy will be required to separate the opposite charges when the process takes place in a suitably oriented electrical field provided by another neighbouring molecule (cf. p. 250). Whilst, therefore, the oxime ether molecules rearrange with difficulty when surrounded solely by non-polar molecules of carbon tetrachloride, the change will take place with increasing ease as the meeting of two strongly polar solute molecules becomes more and more frequent with rising concentration of the solution.

According to this hypothesis, polar molecules of any kind should catalyse the change in the order of the strength of the dipoles in their molecules. In confirmation of this view it was found that when a polar solvent was mixed with the carbon tetrachloride the reaction velocity increased in the manner shown in Figure XII on page 359. The velocity of change could be connected with the concentration, C, of the catalysing solvent by the empirical equation $k = k_{un} + xC + yC^2$, where k_{un} is the velocity in the absence of the catalyst.

A further test of this hypothesis of catalysis by an electrical field was possible, since molecules containing two equal and opposite dipole moments should act as catalysts, provided the two dipoles be far enough apart to give rise to appreciable external polar fields. Both p-dichloro-

benzene and *trans-i*: 4-dibromo*cyclo*hexane were found to have catalytic powers, though both these molecules have a zero resultant dipole moment.

These observations illustrate very clearly the influence of the negative radical on the Beckmann change, and suggest that the essential feature of the process is an intramolecular rearrangement of a kation, as in the Wagner-

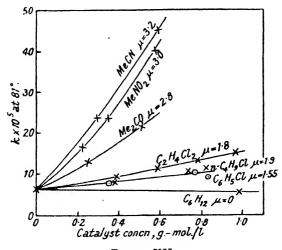


FIGURE XII.
(From A. W. Chapman, J. Chem. Soc., 1934, 1552.)

Meerwein transformation. Mills 1 and Kenner 2 have also pointed out that the Beckmann transformation has certain features in common with the Walden inversion. Thus, if

the kation R_1 , which can be derived equally from

the α - and β -oxime, were to become detached from the system in which transformation is proceeding, the final product would be identical in the two cases, just as the Walden inversion would, under similar conditions, inevi-

¹ B.A. Report, 1932; Nature, 1932, 130, 567.

² Nature, 1932, 130, 309.

tably yield a completely racemized product. In each case, therefore, the migrations of the two interchanging radicals must be pictured as simultaneous and interdependent, rather than as consecutive and independent; and the kation must be regarded as an intermediate phase, which is handed over directly from one anion to another without being allowed even a momentary free existence. A model showing the simultaneous wandering of the hydroxyl and hydrocarbon radicals has been suggested by Mills ¹ as a means of interpreting the *trans*-relationship in the isomeric change of the ketoxime (Fig. XIII), and similar considera-

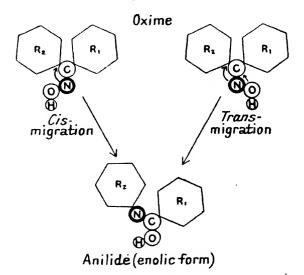


FIGURE XIII.

The rearrangement of the OH group is easily effected during a transmigration, but in a cis-migration it would be impeded by the movement of the N atom.

(By permission of Dr. Mills, from the report of the British Association, 1932.)

tions, based upon a conservation of momentum, are used to account for the *trans*-elimination of water from the aldoximes (Fig. XIV).

¹ Mills, British Association Report, 1932; Nature, 1932, 130, 567.

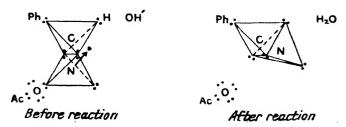


FIGURE XIV.

In the trans-elimination of AcO'H from benzaldoxime acetate the AcO and H ions separate in opposite directions; hence no rotation is set up in the residual molecule of benzonitrile, whilst the N and C atoms are driven together.

(By permission of Dr. Mills, from the report of the British Association, 1932.)

Stereoisomeric Change in the Oximes.1

The conditions under which ketones and aldehydes yield stereoisomeric oximes are of interest in relation to the conditions under which these oximes are converted into isomeric amides. Brady and Dunn 2 pointed out that "there is but one general method by which the aromatic antialdoximes may be converted into their syn-isomerides, namely, through the intermediate formation of their hydrochlorides", although "in a limited number of cases it is possible to convert the anti- into the syn-isomeride by the action of light". They also refer to "the reverse change by which the syn- pass over into the anti-derivatives" as "a reaction which takes place slowly even in the solid state, and more rapidly in solution or in the presence of traces of mineral acids".

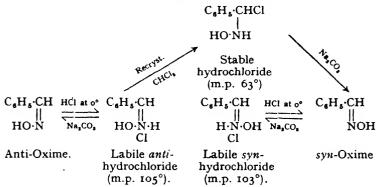
In the special case of benzaldoxime, these changes were attributed by Luxmore 3 to the formation of a labile antihydrochloride (m.p. 105°) which passed into the stable syn-hydrochloride (m.p. 63°) on recrystallizing from chloroform, the relative stability of the oximes being reversed

¹ The configurations of the *syn*- and *anti*-compounds cited under this heading must be reversed in view of the later experiments recorded on p. 353.

² J. Chem. Soc., 1916, 109, 656.

³Ibid., 1896, **69,** 183.

in their salts. Brady and Dunn, however, prepared a third hydrochloride (m.p. 103°) by acting on the syn-aldoxime with hydrogen chloride at a low temperature and concluded that the two hydrochlorides of high melting-point are the anti- and syn- salts of the oxime, and that Luxmore's stable hydrochloride (m.p. 63°) is a compound in which the carbon-nitrogen bond is broken and in which free rotation about the single link is possible.



The mechanism of stereoisomeric change in oximes has been investigated more fully by Taylor and Roberts,1 who studied the (non-reversible) conversion of α - into β -benzilmonoxime. In alcoholic solution, catalysis by hydrogen chloride is inappreciable below molar concentration, but increases very rapidly above that concentration. Lithium, potassium and tetramethylammonium chlorides are more inwerful catalysts than hydrogen chloride, but potassium acetate shows no catalytic effect. "The results suggest that the catalytic agent in these cases is the associated ionpair (or undissociated molecule) of the electrolyte, and that the mechanism is an interaction between the electronic systems of the ion-pair and the carbon-nitrogen double bond." This was further confirmed by showing that hydrogen chloride brought about exceedingly rapid conversion of the α - to the β -oxime in both dry benzene and dry ether solutions.

The difference in stability between the isomeric forms of the oximes and of their hydrochlorides, and the regular change from stability in an oxime to instability in its hydrochloride can be understood if one examines the vector orientations of the dipoles in the two compounds. In the oxime itself both the tervalent nitrogen atom and the hydroxyl group are dipoles, whereas in the kation of the hydrochloride the nitrogen atom is a positive pole, but not a dipole. The resultant dipoles are thus directed as indicated in the diagrams below:

Any dipole due to the hydrocarbon radicals R—C—R' will therefore increase the stability of an oxime but decrease that of its hydrochloride, or vice versa.

The Hofmann and Curtius Degradations.

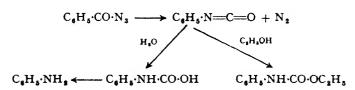
The Hofmann and Curtius degradations provide two alternative methods of converting an acid (through the amide or azide) into an amine. In this process the carbonyl group is eliminated and the number of carbon atoms in the chain is reduced by one, thus providing a method of passing from the C_n to the C_{n-1} series. (a) Stieglitz pointed out in 1896 that these degradations resemble the Beckmann rearrangement in that each reaction involves the migration of a hydrocarbon radical from a carbon atom to a vicinal nitrogen atom. As a mechanism for the two degradations, he suggested that isocyanates were formed as intermediate products, in accordance with the following schemes:

¹ Amer. Chem. J., 1896, **18**, 751; 1903, **23**, 49. ² Ber., 1882, **15**, 407.

364 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

Curtius Degradation.

This mechanism received support from the fact that Hofmann 1 and Tiemann 2 detected the odour of isocyanates in the Hofmann reaction, whilst Lengfield and Stieglitz 3 in 1893 obtained the urethane, R.NH-CO-OEt by carrying out the degradation in alcoholic solution. Moreover, Mauguin in 1909 4 actually obtained methyl isocyanate by heating the dry sodium salt of acetobromamide. Similarly the formation of an intermediate isocyanate in the Curtius degradation has been substantiated by Schroeter 5 and Forster, 6 who showed that the acyl azides give isocyanates and nitrogen quantitatively when heated in dry benzene, and that the products obtained by decomposition in presence of water or alcohol result from secondary reactions of the isocyanate, e.g.,



(b) The degradations, as formulated above, depend on the formation of a compound of univalent nitrogen, $R \cdot CO \cdot N <$. In the Curtius reaction the separation of a molecule of nitrogen could proceed by a non-polar dissociation, of a type which has been discussed in Chapter VII.

Schroeter discovered several analogous reactions on heating di- and triazoles. For example, azobenzil, on heating

¹ Ber., 1882, 15, 407. ² Ber., 1891, 24, 4162.

³ Amer. Chem. J., 1893, 15, 215, 504.

⁴ Compt. rend., 1909, 149, 790.

⁵ Ber., 1909, 42, 2336, 3356; Chem. Zeit., 1908, 32, 933.

⁴ J. Chem. Soc., 1909, 95, 433.

in an indifferent solvent, gives nitrogen and diphenyl ketene, and the reaction was represented as follows:

$$\begin{array}{c|c} C_{\mathfrak{o}}H_{\mathfrak{s}}.C & N \\ \hline C_{\mathfrak{o}}H_{\mathfrak{s}}.C & O \end{array} \xrightarrow{-N_{\mathfrak{s}}} \begin{array}{c} C_{\mathfrak{o}}H_{\mathfrak{s}}.C \\ \hline C_{\mathfrak{o}}H_{\mathfrak{s}}.C & O \end{array} \xrightarrow{C_{\mathfrak{o}}H_{\mathfrak{s}}} C = C = O$$

He therefore followed Stieglitz in formulating the Beckmann transformation as depending on a similar dissociation, involving the formation of a derivative of univalent nitrogen.

nitrogen.

$$R' \longrightarrow C = N \cdot OH \longrightarrow R' \longrightarrow CCl - N \longrightarrow R' \cdot CCl - N \longrightarrow R' \cdot C \longrightarrow R' \cdot$$

This view was criticized by Montagne 2 on the ground that stereoisomeric oximes of the type R'R"C=NOH would yield identical (racemic) products at each of the intermediate stages, and could not therefore give isomeric amides as the final products of the change. He therefore concluded that "the occurrence of an intermediate product containing a univalent atom of nitrogen cannot be the cause of the Beckmann transformation". He also took the opportunity of refuting Wallach's suggestion 3 that an intermediate ring-compound is formed by eliminating a molecule of water from the hydroxyl group of the oxime and a hydrogen atom of the migrating alkyl or aryl group, since this process would convert a p-substituted oxime into a m-substituted amide, in direct contradiction to his own experimental evidence which showed that 4:4'-dichlorobenzophenone oxime gives p-chlorobenzoic acid and pchloroaniline. In the case of the alkyl compounds this point had already been settled by Stieglitz,4 who

¹ Ber. 1909, 42, 2336, 3356; Chem. Zeit., 1908, 32, 933.

² Ibid., 1910, 43, 2014. ⁸ Ibid., 1906, 345, 273. ⁴ Cf. Whitmore and Homeyer, J.A.C.S., 1932, 54, 3438.

366 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

showed that trimethylacetobromoamide gave terbutylamine

$$CMe_3 \cdot CO \cdot NH_2 \longrightarrow CMe_3 \cdot NH_2$$

whereas if an intermediate ring compound were formed the product would be *iso*butylamine:

$$CH_3 \cdot CMe_2 \cdot CO \cdot NHBr \longrightarrow Me_2C \cdot CH_2 \cdot NH \longrightarrow Me_2C \cdot CH_2 \cdot NH_2$$

(c) The criticisms of Montagne do not affect the view that the Curtius degradation might involve a non-polar dissociation of a molecule of nitrogen from the azide; but this does not prove that the Hofmann degradation is also a non-ionic process. On the contrary, the sodium salt of acetobromoamide isolated by Mauguin yields bromide ions as well as sodium ions when decomposed by heat into sodium bromide and methylisocyanate. This process can be formulated most readily as follows:

$$\begin{array}{cccc} CH_3 \cdot CO \cdot NH_3 & \xrightarrow{Br_4} & CH_3 \cdot CO \cdot NHBr & \xrightarrow{NaOEt} & CH_3 \cdot C & \\ & & & & \\ & & & & \\ & & & & \\ & & & & \\ & & & & \\ & & & & \\ & & & & \\ & & &$$

By writing the sodium compound as an oxy-salt, instead of as a nitrogen-derivative, CH₃·CO·NNaBr, the analogy with the ionic mechanism of the Beckmann transformation becomes complete. Indeed, the elimination of bromide ions in the vital stage of the transformation provides even stronger and more direct evidence of an ionic mechanism than in the Beckmann transformation, where the ionic mechanism is based upon the inference that strongly negative radicals tend to react and to migrate in the form of anions rather than as neutral free radicals.

³ Cf. Whitmore, J.A.C.S., 1932, 54, 3274.

¹ Cf. Stieglitz, Amer. Chem. J., 1903, 29, 49.

² Hoogewerff and van Dorp, Rec. trav. chim., 1887, 6, 373; 1889, 8, 173.

Whitmore ¹ has further strengthened the analogy with the Beckmann transformation by pointing out that the migrating hydrocarbon anion cannot have any independent free existence. Thus when active (d)-3:5-dinitro-6- α -naphthylbenzamide (I), in which free rotation of the naphthyl group is hindered by the presence of the nitro and amide groups (cf. pp. 341–346) is treated with bromine and alkali it yields an active amine (II).²

Hence the anion (III) cannot be liberated entirely, for it would racemize immediately. The change must therefore occur *intra*molecularly, so that the migrating radical remains always in close proximity to the remainder of the molecule.³ The reaction differs essentially from the anionotropic changes discussed in the previous chapter,⁴ for these are brought about by the *complete* liberation of an anion, with consequent electron migration in the resulting kation.

(d) The great similarity between the Curtius and the Hofmann degradation reactions leads one to suppose that they both occur by a similar type of molecular rearrangement. The explanation of the Hofmann degradation in

¹ J.A.C.S., 1934, 56, 1427.

² Wallis and Moyer, J.A.C.S., 1933, 55, 2598.

⁸ Whitmore and Fleming, J. Chem. Soc., 1934, 1269.

⁴ Cf. Burton, J. Chem. Soc., 1934, 1268.

the previous section is consistent with our knowledge of the behaviour of ions, but, on the other hand, the nonpolar interpretation of the Curtius reaction is *not* consistent with our knowledge of the behaviour of free radicals.

Firstly, the "univalent nitrogen" structure (A)

$$R \longrightarrow C \bigvee_{N \subset A}^{O} \text{ or } R:C \stackrel{\vdots}{\cdot} \stackrel{\vdots}$$

contains a neutral nitrogen atom possessing only six electrons, two of which are unpaired. It is thus unlike all known free radicals, which possess only one unpaired electron, and, consequently, one would expect that these singlet electrons would interact to form a spin-balanced pair (B), which, however, leaves the nitrogen atom with two lone pairs of electrons.

Secondly, a neutral molecule (A) or (B) would not undergo a polar change, and even if the migration of a radical were to supply the nitrogen atom with the two electrons which it lacks, then one would be (a) producing a charged product from an uncharged one and (b) splitting an electron pair.

Moreover, the Curtius reaction will occur readily in polar solvents such as water or alcohol, and is greatly accelerated by the use of acids.

In view of this difficulty of obtaining a complete non-polar representation of the Curtius rearrangement, one is led to examine whether the change, like that of the Hofmann degradation, might be polar throughout, depending (i) on the formation of a polarized oxy-compound which (ii) degrades with the loss of a negative ion, leaving a residue (iii) which becomes stabilized to an isocyanate by migration of a hydrocarbon anion.

Inspection of the two changes immediately reveals the important fact that in both instances polar activation of the carbonyl group initiates the whole sequence of changes.

So, similarly:

This method of representation indicates a polar fission of the azide ring of three nitrogen atoms, and separation of a dipolar molecule, N=N, of nitrogen.

This mechanism is consistent with that by which Bradley and Robinson 1 represent the reactivity of diazomethane.

In the case of diazoacetic ester this fission of the ring $N = N^+$ is well substantiated by the fact that

hydrogen ion, which would immediately attach itself to C⁻, is a definite catalyst for the elimination of nitrogen. An ion cannot act as a catalyst for a non-polar change.

This ionic mechanism can be used equally well for the interpretation of Schroeter's decomposition of azobenzil

and therefore leaves open to question many of the mechanisms suggested for nitrogen elimination from di- and triazoles.

It does not, however, invalidate the non-polar explanation of the decomposition of azomethane (p. 145).

The Lossen and Stieglitz Rearrangements.

(a) The molecular rearrangement of the hydroxamic acids, and their acyl derivatives, on treatment with alkali, resembles the Hofmann degradation. Isocyanates are formed, which are hydrolysed to primary amines by the action of water, and yield urethanes in alcoholic solutions.

$$R \cdot CO \cdot NH \cdot OH \xrightarrow{-H_2O} R \cdot N = C = O$$

$$R \cdot NH_2 + CO_2$$

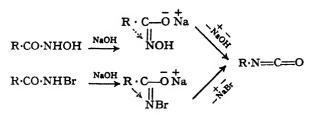
$$R \cdot NH_2 + CO_3$$

$$R \cdot NH_2 + CO_3$$

$$R \cdot NH_2 + CO_3$$

The hydroxamic acids resemble the bromoamides in structure and their molecular rearrangement can be interpreted in the same way, thus:

¹ Lossen, Annalen, 1877, **186**, 1; 1889, **170**, 252; 1894, **281**, 169; Dougherty and Jones, J.A.C.S., 1924, **46**, 1535.



This interpretation is confirmed by the fact that no rearrangement occurs in alkyl or aryl derivatives $R \cdot CO \cdot NBr \cdot R'$ or in the chloro-imino-esters, $R \cdot C(OR')$: NCl, from which it would be necessary to eliminate an alkyl or aryl halide, R'Cl.

(b) Stieglitz 1 observed a migration of hydrocarbon radicals in tertiary hydroxylamines by the action of phosphorus pentachloride. This can be interpreted as depending on the formation of a derivative of univalent nitrogen:

Ph₃C·NH·OH Ph₃C·NH·Cl Ph₃C·N Ph₄C=NPh In the case of the N-methyl derivative this mechanism cannot be used, but the action may be interpreted (like the Beckmann rearrangement) as depending on the reorganization of an unstable kation when an atom of halogen is elim-

inated as an anion.

¹ J.A.C.S., 1914, 36, 272; J.A.C.S., 1916, 38, 2046.

CHAPTER XV

CONJUGATION

Thiele's Theory of Conjugation.

In a famous paper, published in 1899, I. Thiele called attention to the stability of those unsaturated systems which contain two double bonds separated by one single bond. In particular, the unsaturation of these systems was limited almost entirely to the two outside atoms of the chain, whilst the two central atoms were almost completely saturated. In order to explain this phenomenon, he supposed that the tendency of double bonds to undergo addition-reactions was due to residual affinity, and that these residual affinities were capable of mutual saturation whenever the double bonds occurred in alternate positions to one another. The residual affinities were represented by dotted lines added to the ordinary structural formulæ, and their mutual saturation was represented by uniting two of these lines into a loop, thus

C==	=c-	-c=	=c -	→ C=	C	C=	C
•	•	•	•	•	:	:	•
•	•	•	•	•	•	•	•
•	•	•	•	•	••	• •	•

Systems in which this mutual saturation of residual affinity could occur were described as "conjugated systems". Experimental evidence of the behaviour of these systems is set out below.

(a) Bromination of Butadiene. The addition of bromine to butadiene was first examined by Griner, who recorded that a liquid dibromide was first obtained which soon changed into a solid m.p. 53-54°. Thiele showed

¹ Annalen, 1899, 306, 87.

² Compt. rend., 1893, 116, 723; 1893, 117, 553.

that this solid 1 was a 1:4-dibromobutadiene, containing a double bond in the middle of a chain:

$$CH_1 = CH - CH = CH_1 + Br_2 \rightarrow BrCH_2 \cdot CH = CH \cdot CH_2 Br$$

The migration of the bond is reversible, since the dibromide gives with potassium hydroxide an α -bromobutadiene in which the double bonds have returned to their original positions,²

$$BrCH_{\bullet}\cdot CH = CH \cdot CH_{\bullet}Br \xrightarrow{KOH} BrCH = CH - CH = CH_{\bullet}$$

Similarly, with zinc dust and alcohol, butadiene itself is regenerated:

$$BrCH_{2}\cdot CH = CH \cdot CH_{2}Br \xrightarrow{Z_{n}} CH_{2} = CH - CH = CH_{2}$$

More recently, Farmer, Lawrence and Thorpe 3 have verified the original observation of Griner that a liquid dibromide is first formed. This is a I:2-dibromobutadiene, which at 100° changes reversibly to the I:4-dibromide, yielding an equilibrium-mixture containing 20% of the I:2- isomer and 80% of the I:4- isomer.

20% BrCH₂·CHBr·CH=CH₂
$$\rightleftharpoons$$
 BrCH₂·CH=CH·CH₂Br 80% It is therefore an open question whether the 1:4-dibromo compound is really formed as a first product of the interaction or whether it is produced by isomeric change from the less stable 1:2- addition product.

(b) Reduction of Benzil. Thiele suggested that benzil might be regarded as a conjugated system, and that the initial product of reduction should be an unsaturated glycol (I) and not benzoin (II)

¹ Annalen, 1899, 308, 337.

² Willstätter and Bruce, Ber., 1907, 40, 3979.

³ J. Chem. Soc., 1928, 729.

The enolic isomer I could not be isolated, but by reducing benzil in presence of acetic anhydride and sulphuric acid he actually prepared the *cis*- and *trans*- forms of the enolic diacetate. In a similar way Klinger has isolated the dibenzoate by the action of sodium amalgam on benzoyl chloride in moist ether.

$$\begin{array}{c} C_{\mathbf{c}}H_{\mathbf{5}}\cdot \mathrm{CO}\cdot \mathrm{Cl} & \xrightarrow{C_{\mathbf{c}}H_{\mathbf{5}}\cdot \mathrm{C}} = \mathrm{O} & \xrightarrow{C_{\mathbf{c}}H_{\mathbf{5}}\cdot \mathrm{C}\cdot \mathrm{OH}} & \xrightarrow{C_{\mathbf{c}}H_{\mathbf{5}}\cdot \mathrm{C}\cdot \mathrm{O}\cdot \mathrm{Co}\cdot \mathrm{Co}\cdot \mathrm{C}_{\mathbf{c}}H_{\mathbf{5}}} \\ C_{\mathbf{c}}H_{\mathbf{5}}\cdot \mathrm{CO}\cdot \mathrm{Cl} & \xrightarrow{C_{\mathbf{c}}H_{\mathbf{5}}\cdot \mathrm{C}\cdot \mathrm{OH}} & \xrightarrow{C_{\mathbf{c}}H_{\mathbf{5}}\cdot \mathrm{C}\cdot \mathrm{O}\cdot \mathrm{Co}\cdot \mathrm{Co}\cdot \mathrm{C}_{\mathbf{c}}H_{\mathbf{5}}} \end{array}$$

(c) Reactions of Unsaturated Ketones and Acids. The ethylenic linkage —CH=CH— readily unites with hydrogen, whilst the carbonyl group >C=O is specially ready to add on hydrogen cyanide. Thiele pointed out, however, that when the two types of double bond are conjugated with one another in unsaturated ketones containing

add on to the ends of the conjugated chain, rather than to either of the unsaturated groups separately.

$$CH_3:CH:CH=O \xrightarrow{H_3} CH_3:CH:CH:OH \longrightarrow CH_3:CH_2:CHO$$

CH₂:CH·CH=O → CN·CH₂·CH:CH·OH → CN·CH₂·CH₂·CHO
In the same way, the reduction of benzylidene-acetone can
be represented as depending on the initial formation of an
unstable enolic form of the final product:

$$\begin{array}{c} C_6H_5\text{-}CH = CMe = O \longrightarrow C_6H_5\text{-}CH_2\text{-}CH = CMe \cdot OH \longrightarrow \\ C_6H_5\text{-}CH_2\text{-}CO \cdot CH_3. \end{array}$$

This view is strongly supported by the formation of a β -(but never of an α -) compound of the pinacone type as a by-product of this reduction.³

A similar mechanism can be used to explain certain

³ Harries, Annalen, 1897, 296, 295.

curious phenomena observed in the reduction of unsaturated acids, e.g. why $\alpha\beta$ -acids containing conjugated double bonds are easily reduced, but not $\beta\gamma$ -acids.

The reduction of the doubly unsaturated sorbic acid probably proceeds as follows:

$$CH_{3} \cdot CH = CH - CH + CH + COH = COH$$
Sorbic acid.

CH₃·CH=CH—CH₃—CH=C(OH)₁
1:4- addition giving
CH₃·CH=CH—CH₃—CH₃·CO·OH
1:2-dihydrosorbic acid.

CH₃·CH₃·CH=CH—CH—CH=C(OH)₂
1:6- addition giving
CH₃·CH₃·CH=CH·CH₂·CO·OH
1:4-dihydrosorbic acid.

Both I: 4- and I: 6- addition products are formed, but the I: 4- compound preponderates, especially when reduction is effected in alkaline media, e.g. by the action of sodium or of aluminium amalgam.¹

The hypothesis of an initial I: 4- addition to the system C=C-C=O, with subsequent isomeric change, also explains why the ethylenic linkage in unsaturated ketones and acids appears to possess the unusual property of forming addition-compounds with substances such as ammonia, hydrogen cyanide and sodium bisulphite, which usually unite only with the >C=O group.* For example, mesityl oxide with ammonia yields diacetonamine:

$$CH_3$$
 $C=CH-CO-CH_3 + NH_3 \rightarrow CH_3$ $C-CH_3-CO-CH_3$ CH_3 $C-CH_3-CO-CH_3$ CH_3 $C-CH_3$ $C-CH_3$

In the same way several $\alpha\beta$ -unsaturated ketones of the terpene series yield two bisulphite compounds, by 1:4- and 1:2- addition respectively. Addition of hydrogen chloride or bromide can also take place in this manner. Thus, crotonic acid combines with HCl to form β -chlorobutyric

¹ Evans and Farmer, J. Chem. Soc., 1928, 1644; Goldberg and Linstead, J. Chem. Soc., 1928, 2343; Burton and Ingold, J. Chem. Soc., 1929, 2022.

acid exclusively, in direct opposition to the indications of Markownikow's rule (p. 178).

$$\begin{array}{c} \text{CH}_{\bullet}\text{-CH} = \text{CH}_{\bullet} - \text{CH}_{\bullet} - \text{CHCI} = \text{C(OH)}_{\bullet} \rightarrow \\ | & \text{CH}_{\bullet} \cdot \text{CHCI} = \text{CH}_{\bullet} \cdot \text{CO} \cdot \text{OH} \rightarrow \\ \text{OH} & \text{CH}_{\bullet} \cdot \text{CHCI} \cdot \text{CH}_{\bullet} \cdot \text{CO} \cdot \text{OH}. \end{array}$$

(d) The Diels-Alder Reaction. The polymerization of isoprene to dipentene may be cited as an example of a I: 4-addition of an olefine to a conjugated system, thus:

Ring condensation of this type appears to be a general reaction of conjugated systems, for, in 1928, O. Diels, Alder, and their co-workers discovered that butadienes will combine directly with reactive olefines such as acraldehyde, acetaldehyde (as vinyl alcohol), maleic anhydride and p-benzoquinone.¹

¹ Annalen, 1928, 460, 98; 1929, 470, 62; 1930, 478, 139; Ber., 1929, 62, 554, 2081, 2337.

The condensation may occur at room temperature, or on warming the reactants either alone or in an inert solvent. Transient colorations sometimes appear, probably indicating that a radical is added first at one end and then at the other end of the conjugated system. Complex polymerization may sometimes set in; but apart from this experimental difficulty the Diels reaction appears to be diagnostic of a conjugated system, and has been applied widely for the recognition of such systems in many forms of cyclic compounds (cf. Chap. XVI). Its occurrence is dominated by the stereochemical juxtaposition of groups, since the cyclic condensation always leads to the production of a six-membered ring.¹ Thus the cis- and trans- forms of aye-hexatriene both unite with maleic anhydride thus:

Conjugation and Enolization.

Thiele 2 regarded conjugation as an exothermic process, which could be used as a source of energy in promoting chemical change, and even attempted, with the help of thermochemical data and heats of combustion, to estimate the heat-change associated with the development of conjugated systems. The energy difference between conjugated and isomeric unconjugated systems is about 10 cal. per gram molecule, and may be regarded as a sort of heat of conjugation, comparable with a heat of dissolution or with a heat of neutralization of an acid. The conjugated compounds have lower heats of combustion than their unconjugated isomers and are therefore definitely more stable.

In support of this view, many examples could be quoted,

¹ Farmer and Warren, J. Chem. Soc., 1929, 897. ² Annalen, 1899, **306**, 103.

e.g. from the terpene series, of the tendency of double bonds to wander in such a way that conjugated systems are produced, as in the conversion of limonene through terpinolene to terpinene:

Thiele laid special stress on the importance of this factor in promoting enolization. Thus, in accordance with Erlenmeyer's rule, the enolic isomerides of the simple aldehydes, ketones and acids (e.g. vinyl alcohol) are so unstable that only the more stable ketonic forms can be isolated, although their existence is inferred from the possibility of preparing derivatives, such as vinyl chloride, CH₂=CHCl; they have also been assumed to occur in traces as intermediate activated forms in various chemical reactions.

$$\begin{array}{ccc} \text{CH}_3 & \longrightarrow \text{CH}_3 \cdot \text{CHO} & \text{CH}_3 & \longrightarrow \text{CH}_3 \cdot \text{CO} \cdot \text{CH}_3 \\ \text{Vinyl} & \text{Acetalde-} & \text{Acetone.} \\ \text{alcohol.} & \text{hyde.} & \end{array}$$

In the β -diketones and β -ketonic esters, however, a carbonyl group in the β -position has the effect of stabilizing the enolic form, although this does not happen when the carbonyl group is in the contiguous α -position. Thiele pointed out that the enolic form of the β -compounds contains a conjugated system, which is absent, both in the ketonic form of these compounds, and in the hypothetical enolic forms of α - and γ -diketones. Thus, the conversion of ethyl acetoacetate (in the presence of an alkaline catalyst) into an equilibrium mixture with ethyl- β -hydroxycrotonate is

accomplished by the development of a conjugated system, as follows:

In γ -diketones, such as acetonyl-acetone, where two methylene groups are interposed between the carbonyl radicals, the simple enolic form would not be conjugated; but the tendency to conjugation is sufficiently strong to cause the compound to behave as a di-enol, thus:

$$CH_3 \cdot CO \cdot CH_3 \cdot CH_3 \cdot CO \cdot CH_3 \rightarrow CH_3 - C - CH_3 \cdot CH_3 \cdot$$

In view of these and many other similar observations, there can be little doubt that Thiele was right in regarding conjugation as the main factor in promoting the formation of enolic isomerides from carbonyl compounds; in the most stable types of enolic compounds, however, conjugation may be assisted by a process of co-ordination, which favours still further the development of the enolic form. Thus, the enolic form of acetylacetone may be formulated in the same way as its co-ordinated metallic derivatives with a quasi-aromatic ring as follows

Alternation in Conjugated Systems.

Conjugated systems are characterized by a transmission of chemical reactivity from one atom to another. Thus, when ammonia is added to mesityl oxide (p. 375), the NH₂ group does not react with the carbon atom of the >C=O group, as it does in the formation of aldehyde-ammonia, but with a more remote carbon atom, which becomes endowed with an unusual reactivity through being part of the conjugated system. Similarly the proton, instead of

uniting with the oxygen atom in the system C=C-C=O, adds on to a carbon atom, thus,

$$>$$
C $=$ C $=$ O $+$ NH, \rightarrow $>$ C $=$ C $=$ O $=$ NH, H

In the same way, strong experimental evidence can be adduced to suggest that the bromination of butadiene may take place through a I:2 addition of bromine followed by an interchange of bromine and hydrogen in the 2 and 4 positions to yield the I:4-addition product. In each case, however, the fundamental point is an alternation of affinity, whereby alternate atoms acquire similar polarities and similar reactivities. Thus, in the $\alpha\beta$ -unsaturated ketones, the negative polarity of the oxygen, which enables it to unite with a proton, is transferred to the α -carbon atom; and the positive polarity of the carbon in the >CO group is not shared with the adjacent α -carbon atom (which acquires a negative polarity, as we have seen) but jumps over it and is transferred to the β -carbon atom instead.

The important conception of alternation was put forward by Lapworth as long ago as 1898 in a paper on "A Possible Basis of Generalization of Intramolecular Changes in Organic Compounds". It was then suggested that in intramolecular changes the labile group moves from an a-atom to attach itself to a y-atom, or more generally that "a labile group might move along a chain of alternately single and doubly-bound atoms, the ethylenic and single linkings changing places in the path of the labile group".

Lapworth's paper was published in the year preceding that in which Thiele's paper appeared. It deals with similar systems, but in a fundamentally different manner, since Thiele in his main discussion attributed similar "residual affinities" to all the atoms in the conjugated chain (except the terminal atoms, where the residual affinities were unpaired). Lapworth, on the other hand, attributed similar properties only to alternate atoms in the chain.

and so provided a firm foundation for all subsequent work on chemical reactivity in conjugated and aromatic systems.

Electronic Interpretations of Conjugation.

(a) The dotted lines which Thiele superposed on the ordinary bonds in order to represent the "partial valencies" of unsaturated and conjugated systems, cannot be translated directly into symbols which are in accord with the electronic theory of valency. No comparison between normal valencies and residual valencies was attempted by Thiele himself, who appears to have accepted Werner's view that chemical affinity was a force which could be subdivided indefinitely. Thiele's formulæ have, however, been interpreted as if his partial valencies were the equivalent of a single-electron bond. Thus Thiele's formula I for benzene has been regarded as a precursor of J. J. Thomson's electronic formula III, in which the linkage between the atoms of the conjugated aromatic system is formed by three shared electrons.

This interpretation of Thiele's formulæ is, however, only possible in cyclic compounds. Thus, if Thiele's formula IV for butadiene is interpreted in a similar manner by a system of three-electron bonds as in V, it is only necessary to count the electrons in order to discover that the number used is one short of the total number available. A formula which contains the full complement of electrons, such as VI, no longer represents an electrically-neutral, non-polar system, since one of the terminal atoms has a fractional negative charge, — Le, and the other a fractional positive

charge $+\frac{1}{2}c$. It therefore no longer represents Thiele's residual affinities.

R. Kuhn, however, has recently employed formulæ of this type in representing certain of the reactions of conjugated systems with monovalent elements and free radicals, such as sodium or nitric oxide. In support of his theories, to which reference has already been made in an earlier chapter (p. 128). he has pointed out that molecules with long conjugated chains, such as the diphenylpolyenes, (Ph—CH=CH—Ph), 2, or bixin, CH, COO-CH=(CH-CMe=CH), =CH-COOH, exist in only two stereomeric forms, and do not therefore give any indication of cis-trans isomerism about each of the individual double bonds of the conjugated system. 4 Structures with three-electron bonds are therefore not finally excluded from consideration, but they suggest a very different picture to that presented by the partial valency formulæ of Thiele, and their use to explain the common reactions of conjugated systems is very difficult, as can be seen from the very scanty success achieved by W. Hückel in this direction.⁵

(b) A much simpler electronic interpretation of the main facts relating to conjugated systems was proposed by Lowry in 1923,6 in connection with the theory of polar activation of covalent bonds. This interpretation is based upon the hypothesis that only activated molecules undergo chemical change, and that the additive reactions of olefines under the influence of polar reagents depend upon the ionization of one link of the double bond,

^{1 4}th Solvav Report, 1931, cf. p. 128.

² Kuhn and Winterstein, Helv. Chim. Acta, 1928, 11, 87.

³ Karrer, Helfenstein, Widmer and van Itallie, ibid., 1929, 12, 745.

⁴ These facts are easily explicable by the theory of resonance.

⁵ Cf. Theoretische Grundlagen der Organischen Chemie, 1931, vol. I, Chap. 10.

⁶ J. Chem. Soc., 1923, 123, 824.

It then becomes obvious that the characteristic property of a conjugated chain of atoms is its ability to develop a series of alternate positive and negative charges on alternate atoms by the activation of the double bonds.

C=C-C=C becomes
$$\overset{+}{C}$$
- $\overset{-}{C}$ - $\overset{+}{C}$ - $\overset{-}{C}$ - $\overset{+}{C}$ or $\overset{-}{C}$ - $\overset{+}{C}$ - - $\overset{+}{$

In this scheme, the resting form of the molecule is characterized by an alternation of single and double bonds, but in the fully-activated form this is replaced by an alternation of positively- and negatively-charged atoms. It therefore follows that when a conjugated system enters into combination with an ionic reagent, alternate atoms develop ionic charges of like sign, in strict accord with Lapworth's αy -rule, and his more general conception of alternating polarities.

Polar Activation of Conjugated Systems.

(a) In Chapter VIII, the formation of addition-products by unsaturated compounds has been attributed in certain instances to a polar activation of the double bond, usually under the influence of an ionic reagent, giving rise to an unstable semi-polar bond, thus

where arrows have now been added to show that, in the process of activation, a pair of shared electrons is transferred from the double bond to the atom of carbon or of oxygen which acquires a unit negative charge in the activated product. In conjugated systems this process will give rise to activated molecules of the following types:

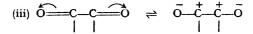
(i)
$$>c$$
 $\stackrel{-}{\sim} c$ $\stackrel{-}{\sim} c$ $\stackrel{-}{\sim} c$ $\stackrel{-}{\sim} c$

(i) In conjugated hydrocarbons, the individual atoms may acquire either a positive or negative charge, exactly as in

the polar activation of ethylene; but it is assumed that consecutive double bonds in the conjugated chain will be ionized in such a way as to produce a regular alternation of positive and negative charges.

(ii)
$$>c$$
 $\stackrel{\leftarrow}{=}c$ $\stackrel{\leftarrow}{=}c$ $\stackrel{\leftarrow}{=}c$ $\stackrel{\leftarrow}{=}c$

(ii) In unsaturated aldehydes and ketones, the sign of the charges will be determined by the tendency for oxygen rather than carbon to acquire a negative charge; the terminal atom of oxygen will thus become negative, whilst the remaining carbon atoms of the conjugated chain will become alternatively positive and negative.



(iii) In the case of orthodiketones, however, the tendency for each atom of oxygen to acquire a negative charge will be likely to overcome the tendency to alternation and a system of crossed polarities may result. The peculiar optical properties which are associated with this type of conjugation are described below (p. 399).

The formulæ thus assigned to the activated forms of butadiene or of crotonaldehyde provide an electrostatic analogue of a chain of magnets, which neutralize one another in the middle of the chain without weakening the terminal poles. Thus in the middle of the system the lines or tubes of force will run between contiguous atoms, and will be almost entirely closed; but a considerable external field will persist round the terminal atoms, thus

This analogy was pointed out by Thiele, who suggested that the atoms of a double bond might be positively and negatively charged and that in a conjugated system the

inner charges of neighbouring double bonds might neutralize one another, thus:

$$\stackrel{\overset{\scriptscriptstyle \leftarrow}{\scriptscriptstyle \leftarrow}}{\overset{\scriptscriptstyle \leftarrow}{\scriptscriptstyle \leftarrow}}\stackrel{\overset{\scriptscriptstyle \leftarrow}{\scriptscriptstyle \leftarrow}}}{\overset{\overset{\scriptscriptstyle \leftarrow}{\scriptscriptstyle \leftarrow}}{\overset{\scriptscriptstyle \leftarrow}{\scriptscriptstyle \leftarrow}}}\stackrel{\overset{\scriptscriptstyle \leftarrow}{\scriptscriptstyle \leftarrow}}{\overset{\scriptscriptstyle \leftarrow}{\scriptscriptstyle \leftarrow}}{\overset{\scriptscriptstyle \leftarrow}{\scriptscriptstyle \leftarrow}}\stackrel{\overset{\scriptscriptstyle \leftarrow}{\scriptscriptstyle \leftarrow}}{\overset{\scriptscriptstyle \leftarrow}{\scriptscriptstyle \leftarrow}}\stackrel{\overset{\scriptscriptstyle \leftarrow}{\scriptscriptstyle \leftarrow}}{\overset{\scriptscriptstyle \leftarrow}{\scriptscriptstyle \leftarrow}}\stackrel{\overset{\scriptscriptstyle }{\scriptscriptstyle \leftarrow}}{\overset{\scriptscriptstyle \leftarrow}{\scriptscriptstyle \leftarrow}}}\overset{\overset{\scriptscriptstyle \leftarrow}{\scriptscriptstyle \leftarrow}}{\overset{\scriptscriptstyle \leftarrow}{\scriptscriptstyle \leftarrow}}\stackrel{\overset{\scriptscriptstyle \leftarrow}{\scriptscriptstyle \leftarrow}}{\overset{\scriptscriptstyle \leftarrow}{\scriptscriptstyle \leftarrow}}\stackrel{\overset{\scriptscriptstyle \leftarrow}{\scriptscriptstyle \leftarrow}}{\overset{\scriptscriptstyle \leftarrow}{\scriptscriptstyle \leftarrow}}\stackrel{\overset{\scriptscriptstyle \leftarrow}{\scriptscriptstyle \leftarrow}}{\overset{\scriptscriptstyle \leftarrow}{\scriptscriptstyle \leftarrow}}\stackrel{\overset{\scriptscriptstyle }{\scriptscriptstyle \leftarrow}}{\overset{\scriptscriptstyle \leftarrow}{\scriptscriptstyle \leftarrow}}}\overset{\overset{\scriptscriptstyle \leftarrow}{\scriptscriptstyle \scriptscriptstyle }}{\overset{\scriptscriptstyle }{\scriptscriptstyle \leftarrow}}{\overset{\scriptscriptstyle \leftarrow}{\scriptscriptstyle \leftarrow}}\stackrel{\overset{\scriptscriptstyle }{\scriptscriptstyle \leftarrow}}}{\overset{\scriptscriptstyle \leftarrow}{\scriptscriptstyle \scriptscriptstyle }}}\overset{\overset{\scriptscriptstyle }{\scriptscriptstyle \leftarrow}}{\overset{\scriptscriptstyle }{\scriptscriptstyle \leftarrow}}}\overset{\overset{\scriptscriptstyle }{\scriptscriptstyle \leftarrow}}{\overset{\scriptscriptstyle \leftarrow}{\scriptscriptstyle \scriptscriptstyle }}}\overset{\overset{\scriptscriptstyle }{\scriptscriptstyle \leftarrow}}{\overset{\scriptscriptstyle }{\scriptscriptstyle \leftarrow}}}\overset{$$

These charges were, however, superposed on the double bonds, instead of being provided by a displacement of the shared electrons of the double bond, as in the formulæ proposed above, in which the alternation of single and double bonds in the normal molecule is replaced in the activated molecule by an alternation of positive and negative charges on a chain of single bonds.

The formulæ proposed above can be used, like Thiele's " partial valency " formulæ, to account for the I: 4- addition of a reagent to a conjugated system. Thus the proton H of an ionizable reagent HX will be attracted exclusively by the negatively-charged atom, whilst the anion A will be attracted exclusively by the positively-charged atoms. therefore, a proton H is added to the negatively-charged oxygen atom I of an unsaturated aldehyde, the negative radical A must add on at the 2, 4, 6, etc. position, in accordance with Lapworth's ay-rule. If the system is activated throughout, the attraction will be strongest at the ends of the chain, where the shielding effect of adjacent charges is at a minimum, and addition will take place there. the terminal charges have thus been neutralized by the ions of the reagent, the intermediate charges will be free to neutralize one another and thus give rise to a normal molecule, in which the double bonds occupy the positions formerly occupied by single bonds in the conjugated chain and conversely: thus

On this basis, I: 2- addition to a conjugated system which may amount to as much as 96% in I: 4-diphenylbutadiene, C₆H₆·CH:CH·CH:CH·C₆H₅, can be regarded as the result of interaction with a partly activated molecule, e.g.

>C—C—C—C—C

Vicinal addition in a conjugated system would then result from local activation of a double bond, whilst terminal addition would indicate a spread of polar activation through the whole system.

(b) On the other hand, Lapworth ¹ and Robinson regard it as improbable that any one molecule should be activated simultaneously at more than one point at any one instant, and therefore dissent from depicting activated butadiene as CH₂—CH—CH—CH₂, with both double bonds activated. As an alternative, it may be supposed that the preliminary activation of one bond is followed by the migration of one of the ionic charges thus developed from one atom to an alternate one, together with the requisite redistribution of the covalencies, thus:

Curved arrows are used to indicate the direction of movement of the electron pairs. Thus in the process of electron migration the lone pair of electrons (x) on the original anion leaves an atom of carbon in order to form the second covalent link of the new double bond; consecutively, a pair of shared electrons (y) leaves the double bond in order to form a lone pair on the new anion. In longer conjugated chains, in which an ε migration of ionic charges is possible, a similar diagrammatic representation may be used

A curved arrow (z) placed around an atom, indicates the ¹ J. Chem. Soc., 1922, 121, 417; cf. ibid., 1901, 79, 1265.

shift of a double bond along the chain without the formation of an ionically-charged atom.

The difference between these two methods of representation is of no real importance. The primary process is probably a collision between an ion (or ionizable molecule) of the reagent and one atom of a double bond, giving rise to a local ionization of the bond, exactly as in the case of a simple olefine,

The complex thus formed may either (i) dissociate again or

(ii) collide with an anion \overline{X} (or with a molecule which can donate this ion). If this collision occurs at the contiguous atom of carbon, vicinal addition will occur, and there is no reason why the rest of the molecule should be activated. If, however, the collision occurs at a more distant point of the conjugated chain, an electronic change may proceed throughout the intermediate chain of single and double bonds, as a result of which the ionic charges of the reagent will be neutralized, exactly as in a prototropic change (Chap. XIII).

This mechanism also makes it possible for the radical A to migrate between the 2, 4, 6, etc. positions by a process of reversible isomeric change (i.e. anionotropy).

The essential characteristic of conjugated systems is indeed precisely this possibility of neutralizing the opposite charges of an ionized reagent at the terminal points of a chain of atoms. This feature of conjugated systems has been described as "graphitic conduction" by Lowry, is since

¹ Nature, 1925, 115, 376.

graphite has many of the properties of a conjugated system extended indefinitely in two dimensions.

C. K. and E. H. Ingold 1 in 1926 termed the process of consecutive electron displacement

(e.g.
$$c = c - c = c - c = c - c = c - c = c - \bar{c}$$
)

electromeric change, though even in 1921 H. S. Fry 2 had made use of the term electromerism to denote electron movement.

The designation "electromeric change" is used by Robinson to denote either the polar activation process

or the bond-shifting process

$$c = c - c$$
 i.e. $c = c - c \rightarrow c - c = c$

in each of which there occurs movement of a pair of electrons.

C. K. Ingold in 1926 ascribed electromeric change to the action of a tautomeric effect, and introduced the symbol T to denote the process of shifting an electron pair, for which Robinson had already (loc. cit.) used the curved-arrow symbol ...

Classification of Conjugated Systems.

The preceding discussion shows that the essential characteristic of conjugated systems consists in their ability to undergo a rearrangement of bonds whereby a valency electron can be handed on from end to end of the conjugated chain. In this way provision can be made for neutralizing the opposite charges of an ionized reagent in the process

¹ J. Chem. Soc., 1926, 1313; cf. ibid., 1933, 1121.

² H. S. Fry, "The Electronic Conception of Valency", 1921.

^{3 4}th Solvay Report, 1931: Institute of Chemistry Lectures, 1932; J. Chem. Soc., 1926, 402.

⁴ Chemical Society Annual Reports, 1926, 23, p. 140; cf., however, Robinson, Solvay Reports, 1931.

of "terminal addition" as well as in the more normal process of "vicinal addition"; and the same characteristic provides a mechanism whereby "vicinal addition" can be converted into "terminal addition" by a process of isomeric change, in which a positive or negative radical migrates between alternate atoms of the chain in accordance with Lapworth's ay-rule. In Thiele's classification conjugated systems were confined to those containing olefinic >C-C< and carbonyl >C=O groups; but the work of Decker 1 and of Robinson 2 has shown that any element (e.g. nitrogen, sulphur, etc.) which can form a double covalent bond, or which possesses unshared electrons, can form part of a conjugated system. The reactions of these conjugated systems depend upon the electrical characteristics of the elements present, and can be classified into well-defined groups 3 as follows:

(a) Poly-enoid Systems. —C=C—C=C

Conjugated olefines, such as butadiene, isoprene and aromatic hydrocarbons, may form addition-compounds by direct addition of the ions of the reagent in the I:2, I:4, etc., positions; or these ions may be initially attached to two adjacent atoms of carbon and then migrate to alternate atoms of the conjugated chain. In a symmetrical compound, such as butadiene itself, there is of course no difference between initial addition at one or other end of the conjugated chain; but in unsymmetrical compounds it will be more easy to activate one of the double bonds than the others. For example, phenylbutadiene brominates almost entirely in the 3:4 position, that is to say, at the double bond which is most distant from the phenyl group.

 $C_6H_5\cdot CH = CH = CH_2 \rightarrow C_6H_5 - CH = CH - CHBr - CH_2Br$

Effects such as this can be attributed, as in the case of an

¹ Ber., 1904, 37, 523; 1905, 38, 2893.

² J. Chem. Soc., 1916, 109, 1029.

³ Robinson, Solvay Reports, 1931. Institute of Chemistry Lectures, 1932, pp. 19-31.

olefine, to the polarity of the substituents, and are most strongly marked when atoms other than carbon are adjacent to one of the double bonds.

(b) Hetero-enoid Systems, e.g. N—C—C—, O—C—C—.

"Hetero-enoid systems" ² are characteristic of olefines in which the anionoid character of unsaturated carbon is intensified by linking it to a "negative" atom (oxygen, nitrogen, sulphur, etc.) by a single bond which can be converted into a double bond by a rearrangement of valency electrons. Thus Robinson showed in 1916 ³ that the tervalent nitrogen atom of ethyl β -amino-crotonate was apparently conjugated to the rest of the system, since the radicals of methyl iodide could be added to alternate atoms of carbon and nitrogen in place of the usual addition of both radicals to nitrogen:

$$\begin{array}{c} CH_3 & CH_3 \\ NH_4 - C - CH - COOEt + CH_3I \rightarrow \stackrel{+}{N}H_2 - C - CH - COOEt \\ \stackrel{-}{I} & CH_2 \end{array}$$

The tertiary base, $(C_2H_5)_2N$ — $C(CH_3)=CH$ —COOEt, reacted in the same way. This I:3 addition in "crotenoid systems" is only possible because tervalent nitrogen carries a lone pair of electrons, which can be used to form a double bond between nitrogen and carbon.

$$>$$
N \rightarrow c \rightarrow \vec{c} \rightarrow \rightarrow \vec{n} =c \rightarrow \vec{c} \rightarrow

The formation of this double bond depends, in turn, on the rupture of one link of the original olefinic double bond, by transferring a pair of shared electrons to the terminal carbon atom. As a result of these changes the nitrogen

¹Cf. Lucas, J.A.C.S., 1924, 46, 2475; 1925, 47, 1459, 1462; Kermack and Robinson, J. Chem. Soc., 1922, 121, 427; Allan, Oxford, Robinson and Smith, J. Chem. Soc., 1926, 401.

² Robinson, 4th Solvay Report, 1931.

³ J. Chem. Soc., 1916, 109, 1038.

acquires a positive charge, which enables it to annex an iodide ion, whilst the terminal atom of carbon acquires a negative charge and a lone pair of electrons, which enable it to form a covalent bond with the methyl radical CH₃ of the iodide.

The oxygen in enols and in enol ethers also possesses a mobile electron pair and can therefore act in the same manner, e.g. in the alkylation of aceto-acetic ester. This reaction is carried out with the use of sodium (or sodium ethoxide) to produce the reactive enolic form of the ester, but it gives C-substituted derivatives, thus:

$$\stackrel{+}{\text{Na}}\stackrel{-}{\text{O}} \stackrel{-}{\text{C}} \stackrel{-}{\overset{-}{\text{C}}} \stackrel{-}{\overset{-}{\underset{}{\text{CH}_{3}}}} \stackrel{+}{\text{Na}} \stackrel{-}{\text{O}} \stackrel{-}{\overset{-}{\text{C}}} \stackrel{-}{\overset{-}{\underset{}{\text{C}}}} \stackrel{-}{\underset{}{\text{CH}_{3}}}$$

On the other hand, vinyl chloride, Cl·CH—CH₂, is not reactive, for example, towards alkalis, since electron transference of the type indicated above would tend to produce a positively-charged chlorine atom, instead of the negatively-charged atom which is the normal product of hydrolysis by alkalis. The acidity of phenols and the feeble basicity of the aromatic amines, both of which are accompanied by strong reactivity in the hydrocarbon ring, may be explained on similar lines (p. 442).

The extent to which this electromeric change is liable to occur depends, of course, upon the nature of the heteroatom, and to some extent upon the other groups which may be connected to it. The order of effectiveness, which is in reality the order of mobility of the lone pair of electrons, is given by Kermack and Robinson 1 as follows:

This order is identical with that of diminishing strength of bases.

¹ J. Chem. Soc., 1922, 121, 433.

(c) Katio-enoid Systems, e.g. C=C-C=0, C=C-C=N.

These systems contain hetero-atoms which tend to abstract electrons, and thus to produce carbon kations; they are therefore the reverse of the systems considered in the preceding paragraph. $\alpha\beta$ -Unsaturated aldehydes, such as crotonic aldehyde, belong to this class, which was therefore originally described as "crotonoid". The rearrangement of bonds (or of valency electrons) in the katio-enoid systems takes place as follows:

$$>c = c - c = \delta \rightarrow > c - c = c - \delta$$

It is conditioned by the fact that the hetero-atom forms a more stable anion than carbon. The extent to which this electron-transfer will tend to take place therefore follows the order of stability of these anions, and is approximately as follows:

-Cho > -C
$$\stackrel{\mathrm{O}}{\stackrel{\mathrm{C}}{\stackrel{\mathrm{C}}{\stackrel{\mathrm{O}}{\stackrel{\mathrm{C}}{\stackrel{\mathrm{C}}{\stackrel{\mathrm{O}}{\stackrel{\mathrm{C}}{\stackrel{\mathrm{C}}{\stackrel{\mathrm{O}}{\stackrel{\mathrm{C}}}{\stackrel{\mathrm{C}}{\stackrel{\mathrm{C}}{\stackrel{\mathrm{C}}{\stackrel{\mathrm{C}}}{\stackrel{\mathrm{C}}{\stackrel{\mathrm{C}}}{\stackrel{\mathrm{C}}{\stackrel{\mathrm{C}}{\stackrel{\mathrm{C}}{\stackrel{\mathrm{C}}{\stackrel{\mathrm{C}}}{\stackrel{\mathrm{C}}{\stackrel{\mathrm{C}}}{\stackrel{\mathrm{C}}{\stackrel{\mathrm{C}}}{\stackrel{\mathrm{C}}{\stackrel{\mathrm{C}}}{\stackrel{\mathrm{C}}{\stackrel{\mathrm{C}}}{\stackrel{\mathrm{C}}}{\stackrel{\mathrm{C}}}{\stackrel{\mathrm{C}}}{\stackrel{\mathrm{C}}}{\stackrel{\mathrm{C}}}{\stackrel{\mathrm{C}}}{\stackrel{\mathrm{C}}}{\stackrel{\mathrm{C}}}{\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}{\stackrel{\mathrm{C}}}}\stackrel{\mathrm{C}}{\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}{\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}{\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}\stackrel{\mathrm{C}}}\stackrel{\mathrm{C}}$$

(d) Neutralized Systems, e.g. N—C=O.

These systems contain both an electron-donor and an electron-acceptor, which are capable of interaction so that each is stabilized by the other, e.g.

$$NR_z \longrightarrow C \longrightarrow 0 \longrightarrow NR_z \longrightarrow C \longrightarrow \bar{0}$$

The reactivity of both ends of the conjugated system will be diminished by the electromerization, but its occurrence may be betrayed by the production of a molecular dipole moment. The carboxyl group belongs to this type, since a state of dynamic isomerism can occur between the >C=O and the >C-O-H groups. The order of reactivity of analogous groups has already been given (p. 391), but could be amplified to include other atoms.

¹ Allan, Oxford, Robinson and Smith, J. Chem. Soc., 1926, 410.

(e) Discordant (or Crossed) Polar Systems.

These include compounds such as the quinones, the peroxides, and many others, in which two groups are present which tend to promote electromeric changes in opposed directions. They are in consequence unstable, highly reactive substances (cf. p. 384). A typical case is that of maleic acid, which Thiele formulated as a stable conjugated compound, but in which the crossed polarities of the two carbonyl groups produces a state of instability in the double bond.

This instability can sometimes be relieved by isomeric change, as in the conversion of citraconic acid to itaconic acid with strong alkali:

In the same way, the ready conversion of benzil to benzilic acid by the action of alkali (p. 349) may be attributed to the instability caused by the crossed polarities of the —CO—CO— group in the yellow α-diketone.

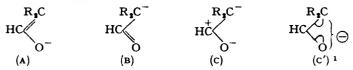
The Stability of Conjugated Systems.

It has already been mentioned (p. 377) that Thiele regarded conjugation as conducive to molecular stability, and estimated that the heat of conjugation of two double bonds was about 10 cal. per gram-molecule. Ingold 1 has recently developed this theme, and, by introducing the conceptions of wave-mechanics, according to which an electron has no exact location, so that its charge may conceivably be shared over many atoms, has represented many conjugated systems as entities in resonance (p. 305), having a "distributed charge". Thus he has suggested that since compounds of the type

$$R_sC \longrightarrow CH \longrightarrow R_sC(H) \longrightarrow CH \longrightarrow O$$
 very readily undergo isomeric change, an ion such as (C)

¹ J. Chem. Soc., 1933, 1125; Chemical Reviews, 1934, 15, 225. Cf. the appendix in J. W. Baker's "Tautomerism" (London, 1934).

or (C'), which has a distributed charge, may possibly have a lower energy content than either of the ions (A) or (B).



(A) and (B) are termed "unperturbed" and (C') the "perturbed" forms of the ion. Lowry had, in 1923, suggested that the symmetry of an activated ion such as (C) might be conducive to its formation during the process of an isomeric change, and pointed out that its external electrical field would be small.

The tendency for an unperturbed molecular structure to lose energy by passing to a perturbed structure, containing a "distributed change", as in (C'), is thought by Ingold to be the driving force that initiates electromeric change in a conjugated system (i.e. initiates the spread of polar activation from one bond to another). It can be likened in many ways to an electrostatic inductive effect (compare p. 239), but, as it can be considered to cause an energy change by the quite different process of coupling together different electronic orbits, Ingold has termed it a mesomeric effect.³ The resulting molecule, in which the orbital vibrations of the different electrons are coupled together, is said to be in a state of resonance.⁴

It must be noted that the mesomeric effect is essentially a stabilizing effect. Thus one can suppose that electromeric change in a conjugated system follows a polar activation, because the ion initially produced has a higher energy content than the ion produced by the subsequent electromerization. Hence, for example, Ingold's conception of energy coupling of electronic orbits provides an explanation of why the activation of a carbonyl group in an $\alpha\beta$ -unsatu-

¹ Compare the structural formula of p. 314.

² J. Chem. Soc., 1923, 123, 822, 2111. 8 Loc. cit.

⁴ Pauling and Wheland, J. Chem. Physics, 1933, 1, 362; Pauling and Sherman, ibid., 1933, 1, 606; Sutton, Trans. Faraday Soc., 1934, 30, 789.

rated ketone should be followed by the activation of the double bond, without a further increment of energy being required, i.e. why C=C-C-O promptly gives C-C-C-O (or C-C-C-O). Physical evidence has been forth-coming to indicate that, owing to the mutual perturbation of electronic orbits, a molecule may indeed acquire a "resonance structure" which cannot be depicted exactly by any conventional graphic formula. Thus Pauling 1 and his collaborators have shown that, in many molecules, the actual heat of formation is less than that calculated for any simple structure from heats of rupture of simple covalent links. Spectroscopic measurements of molecular energy-levels in the infra-red region lend further support to this conclusion.²

The perturbation of electronic orbits will undoubtedly affect the value of the resultant dipole moment of a molecule. Hence Sutton,³ Zahn ⁴ and others ⁵ have ascribed some of the observed deviations from the rule of vector addition of dipole moments (p. 76) to effects of "resonance", and have pointed out that slight, but yet measurable, changes in stereochemical configuration may result, and will have the effect of increasing the geometrical symmetry of a molecule.

The conception of perturbation of electronic orbits can be applied to explain the great stability of many neutralized conjugated systems. Thus the chemical properties of the carboxyl group, and also the magnitude of its dipole moment, can be well represented by the symmetrical formula

¹ J. Chem. Physics., 1933, 1, 606; J.A.C.S., 1931, 53, 3231; 1932, 54, 988.

³ Compare the Discussion on Dipole Moments in *Trans. Faraday Soc.*, 1934, 30, 789-830.

^{*} Ibid., 1934, 30, 856; see also ref. * above.

⁶ Marsden and Sutton, J. Chem. Soc., 1936, 1383; see also ref. ⁴ above.

which may be regarded as a resonance form of the unsymmetrical structure

in which electron movement indicated by the curved arrows would be expected to occur. Resonance is also probable in esters and in amides.

Similarly, in the case of the nitro-group,

$$-\stackrel{+}{N} \stackrel{\circ}{\sqrt{0}}$$
,

one cannot differentiate in any way between the two oxygen atoms, since the large dipole moment of the nitro-group is vectorially directed along their medial line ¹

$$-\overset{+}{N} \overset{O}{\overset{+}{\circ}} \longrightarrow$$

Ingold therefore assigns to it the symmetrical perturbed structure

$$-\frac{1}{N} \left\langle \begin{array}{c} O \\ O \end{array} \right\rangle \ominus$$

with a distributed negative charge.

Sidgwick has utilized the same conception in order to explain the structure of the organic azides, which must contain three nitrogen atoms in a nearly linear arrangement (cf. p. 113).

The stability of aromatic ring compounds (Chap. XVI) and of free radicals such as triphenylmethyl ³ can again be explained most easily by utilizing the idea of resonance.

¹ J. W. Williams, Phys. Zeit., 1928, 29, 174, 683.

² Trans. Faraday Soc., 1934, 30, 802.

³ Burton and Ingold, *Proc. Leeds Phil. Soc.*, 1929, 1, 421; Rodebush, *Trans. Faraday Soc.*, 1934, 30, 821; E. Hückel, *ibid.*, 1934, 30, 40; Ingold, *ibid.*, 1934, 30, 52.

Optical Properties of Conjugated Systems.

Conjugated systems exhibit characteristic optical properties.

(a) Absorption. (i) Saturated hydrocarbons are almost completely transparent in the more easily accessible ultraviolet region, e.g. up to the limit of transmission of light by air and by quartz at about 1850 A.U. and also for a considerable part of the Schumann region; but they appear to possess a very intense absorption in the Schumann region, since their refractive dispersion is dominated by a characteristic frequency at about 1100 A.U., which can be attributed to the shared electrons of the single bonds C—C and C—H. (ii) Unsaturated hydrocarbons, show, in addition, a strong absorption band at the extreme edge of the accessible ultra-violet region with a maximum

addition, a strong absorption band at the extreme edge of accessible ultra-violet region, with a maximum, $\log \varepsilon = 3.7$ at about 1850 A.U., and a step-out (which is 10,000 times less intense) between 2300 and 2800 A.U. (Fig. XV (2)).1 This additional absorption can be attributed to the shared electrons of the double bond. (iii) Conjugated hydrocarbons exhibit absorption bands of a new type, which are of greater intensity and at lower frequency than those of unconjugated hydrocarbons. Thus cyclohexadiene resembles the olefines in showing a strong absorption band with a maximum, $\log \varepsilon = 4.7$, on the edge of the Schumann region at about 1850 A.U.; but it also shows a new absorption band, which has a pair of twin maxima, $\log \varepsilon = 4.0$ at 2560 and 3.8 at 2680 A.U. (Fig. XV (1)).* In hydrocarbons containing two or three conjugated double bonds, the absorption bands are confined to the ultra-violet: but unsaturated hydrocarbons containing more than four conjugated double bonds may give rise to visible colour.3 This absorption is a characteristic attribute of the shared electrons of the conjugated system, but the detailed electronic structure of the system is not yet known.

¹ Allsopp, Proc. Roy. Soc., 1934, A 143, 624.

³ Kuhn and Winterstein, Helv. Chim. Acta, 1929, 12, 899.

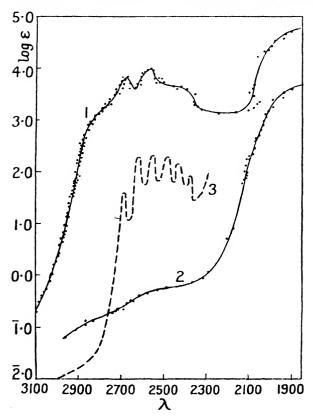


FIGURE XV.1—Molecular extinction coefficients of (1) 1:3-cyclo-hexadiene; (2) cyclohexene; (3) benzene.

(By permission of The Royal Society.)

(iv) Unsaturated Ketones. Camphor shows a characteristic "ketonic" absorption band with a maximum, $\log \varepsilon = 1.57$ at 2880 A.U. in alcohol. This band, which is characteristic of the carbonyl group, is of much lower frequency than that of the olefines. It is attributed to an activation of the shared electrons of the double bond between carbon and oxygen, >C=O. When the ketonic group is

¹ From Allsopp, Proc. Roy. Soc., 1934, A 143, 624.

conjugated with olefinic double bonds in methylenecamphor or in benzylidenecamphor, no new absorption band can be detected; but the intensity of the ketonic band is increased enormously. Thus in benzylidenecamphor the intensity is increased 400-fold to a maximum, $\log \varepsilon = 4.32$ at 2900 A.U.¹; but the wave-length is practically the same as in camphor.

In these two compounds, the polarities of the activated molecules can alternate in a perfectly regular way, starting from the negative oxygen of the ketonic group, and will therefore be "concordant" throughout. In camphorquinone, where the polarities are "discordant" or "crossed", a totally different type of conjugation is disclosed by the absorption spectrum, since the intensity of the band remains practically constant at $\log \varepsilon = 1.46$, but the maximum is displaced by nearly 1800 A.U. to 4650 A.U., where it falls within the blue region of the spectrum and gives rise to a visible yellow colour.

(b) Molecular Refractions. (i) The refractive dispersion of a saturated hydrocarbon, e.g. cyclohexane, C₆H₁₂, depends on the shared electrons between carbon and carbon and between carbon and hydrogen, since the unshared electrons of the inner K shell of carbon are too firmly held to

¹ Lowry and French, J. Chem. Soc., 1924, 125, 1921.

² Lowry and Southgate, J. Chem. Soc., 1910, 97, 905.

⁸ This colour is characteristic of the quinones, but the olefinic double bonds are only required in order to link up the carbonyl groups of the p-quinones.

influence the refractive power in the visible or the ultraviolet region. The refractive dispersion can be expressed by a constant term, plus one variable term with a characteristic frequency in the Schumann region at about 1100 A.U. This frequency is presumably a weighted mean of the characteristic frequencies of the two types of shared electrons, which appear to have energies of activation of the same order of magnitude. Empirically, the molecular refraction

$$(n-1)\frac{M}{d}$$
 (Gladstone and Dale) or $\frac{n^2-1}{n^2+2}\frac{M}{d}$ (Lorentz and Lorenz)

can be expressed in simple cases of this sort as the sum of a series of atomic refractions

This is another way of expressing the fact that the molecular refractions are an additive function of the partial refractions of the different types of outer electrons, including the unshared electrons of the halogens as well as the shared electrons which link the halogen to the carbon skeleton.³

(ii) The refractive dispersion of an unsaturated hydrocarbon, e.g. cyclohexene, contains an additional term with a characteristic frequency in the ultra-violet, but this term is so weak that it can scarcely be detected outside the region of absorption. Empirically it has been found that the effect of "isolated double bonds" can be represented by adding to the molecular refraction a constant increment for each double bond, e.g. 2.4 for H 6563 (Gladstone and Dale), 1.733 for Na 5893 (Lorentz and Lorenz). This increment, like the additional partial refraction in the dispersion equation, must be associated with the shared electrons of the double bond; but it is not yet possible to develop any direct relationship between these two quantities, apparently because the partial refractions of the other

¹ Landolt, Annalen, 1882, 213, 75.

⁴ Allsopp, Proc. Roy. Soc., 1934, A 143, 624.

single bonds are enhanced when one of them is replaced by a double bond. The increment is, however, a constant quantity for double bonds which are either contiguous ("cumulative") as in allene CH₂—C—CH₂, or are separated by more than one single bond.

(iii) Brühl pointed out in 1907 that conjugated systems, containing double bonds separated by one single bond, exhibit the phenomenon of optical exaltation, i.e. the molecular refraction is higher than the value calculated for two double bonds or than that observed in isomeric compounds containing two "isolated" double bonds,

```
e.g.  CH_{\mathfrak{g}} = CH - CH_{\mathfrak{g}} - CH_{\mathfrak{g}} - CH = CH_{\mathfrak{g}}   CH_{\mathfrak{g}} = CH - CH_{\mathfrak{g}} - CH = CH_{\mathfrak{g}}   CH_{\mathfrak{g}} - CH = CH - CH_{\mathfrak{g}} - CH = CH_{\mathfrak{g}} - CH_{\mathfrak{g}}   2 : 4 - Hexadiene   conjugated `` double bonds.   M_{\mathfrak{g}} = 28 \cdot 97 \text{ (calc.)}   M_{\mathfrak{g}} = 28 \cdot 97 \text{ (calc.)}   M_{\mathfrak{g}} = 28 \cdot 97 \text{ (calc.)}   Exaltation = 1 \cdot 67
```

This phenomenon, however, is not universal, especially in ring-compounds;

e.g. Benzene. Toluene.
$$M_D = 26 \cdot 19 \text{ (obs.)} \qquad M_{\alpha} = 30 \cdot 79 \text{ (obs.)}$$

$$M_D = 26 \cdot 30 \text{ (calc.)} \qquad M_{\alpha} = 30 \cdot 89 \text{ (calc.)}$$

$$Diff. - 0 \cdot 11 \qquad Diff. - 0 \cdot 10$$

The normal value of the molecular refraction of benzene was used at an early date as an argument in favour of Kekulé's formula, but was subsequently explained away on the ground that the symmetrical distribution of the double bonds gave rise to neutral conjugation instead of to actual conjugation, as in open chain systems. This conclusion that the formation of a closed circuit of alternate single and double bonds suppresses the usual effects of conjugation was confirmed by the discovery of Willstätter and Waser in 1911 that cyclooctatetrene,

¹ Ber., 1907, 40, 878 and 1153; J. Chem. Soc., 1907, 91, 117.

² Brühl, Annalen, 1880, 200, 229.

³ Brühl, Ber., 1907, **40**, 887.
⁴ Ber., 1911, **44**, 3423.

which contains four pairs of double bonds in a closed ring, shows the same absence of optical exaltation as benzene itself:

CH—CH—CH—CH

CH—CH—CH—CH

$$M_D = 35 \cdot 20 \text{ (obs.)}$$
 $M_D = 35 \cdot 24 \text{ (calc.)}$

Diff. — 0.04

Brühl predicted that, in the absence of a closed circuit of single and double bonds, optical exaltation would occur in I: 3-cyclohexadiene, where the double bonds are conjugated, but not in the I: 4-isomer, where the double bonds are on opposite sides of the aromatic ring, and this prediction has been widely accepted as if it were an experimental fact. Actually four typical cyclic dienes have given the following results:

In all these compounds the conjugation forms an "open circuit", but the conjugated system, instead of producing exaltations of about 1.6 units, gives increments of -0.43, +0.07, +0.89 and -0.17 in the four hydrocarbons. On

the other hand, a very large exaltation is observed in the hydrocarbon

which gives an exaltation of 2.49, exactly as in hexatriene, CH₂=CH-CH=CH-CH=CH₂. It therefore appears likely that, whilst many factors (including the presence or absence of a side-chain) may influence the development of optical exaltation, the occurrence of a parallel configuration of double bonds may be a very important factor in promoting the optical exaltation of conjugated hydrocarbons in the liquid state, as it certainly is in developing abnormal refractive power in crystalline compounds containing systems of this kind.

From the physical point of view the conjugation of two double bonds in cyclohexadiene is associated with the development of an additional partial refraction, with a characteristic frequency in the ultra-violet, but this is much stronger than in cyclohexene, and contributes a definite increment to the refraction even in the visible spectrum. This increment, which must be associated with the conjugated double bonds, does not, however, produce any marked optical exaltation in the visible spectrum, where it appears to be balanced by decrements in the partial refractions of the single bonds; but in the ultra-violet, as the absorption bands are approached, it gives rise to very large increments, and to refractive indices which are finally larger than those of benzene, with three double bonds instead of only two.

The chemical significance of molecular refraction has already received discussion in Chapter XI, in which a connection with molecular polarizability was indicated. On general grounds one would anticipate that the polarizability of a covalent bond would be, in addition, a measure of its ease of polar activation, but it does not follow that the responses of the valency electrons to the large electrical disturbances which can rupture a bond should be propor-

tional to their responses to the much smaller alternating electrical fields produced by the passage of light waves through a molecule. Hence one must proceed with great caution in correlating optical properties with chemical reactivity, particularly when dealing with complicated molecules containing conjugated systems within which intramolecular actions and reactions are very facile.

CHAPTER XVI

AROMATIC COMPOUNDS-I

A. AROMATIC STRUCTURE

It has long been customary for chemists to differentiate sharply between "aromatic" and "aliphatic" compounds, when considering both their structures and their chemical reactivities, but nevertheless it has been realized that the differences between the two groups of substances must be of degree rather than of kind. The main peculiarities of aromatic character can be indicated from the statement that whereas it would be anticipated, from their very low percentage of hydrogen, that substances like benzene, naphthalene, pyridine and thiophene would be unstable and extremely unsaturated compounds, prone to undergo instantaneous oxidation and addition of halogens, yet they are in fact extremely stable substances which, like the paraffin hydrocarbons, react most easily by substitution processes.

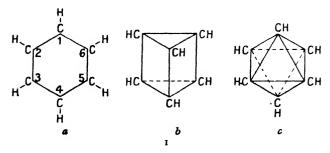
(a) The Kekulé Formula.

"To Kekulé must be assigned the credit of having provided the first key to the elucidation of the nature of aromatic compounds, since, in 1865, he suggested that benzene contained a symmetrical ring of six carbon atoms, each having attached to it one hydrogen atom. Important investigations, amongst which those of Ladenburg are notable, carried out very soon afterwards, established the

¹ Bull., 1865, 3, 98. Cf. "Chemie der Benzolderivate", Erlangen, 1867.

² Cf. Meyer-Jacobsen, "Lehrbuch der Organischen Chemie", Vol. II, 1, 46-84.

fact that benzene must be represented by a *planar* ring structure, I (a), rather than by a prism or octahedral structure, (b) or (c), since only in this way is it possible to account



satisfactorily for the isomerism and interrelationships of the substitution derivatives of benzene, and also for absence of any molecular dissymmetry in them.

The skeleton formula of a plane hexagon (Ia) even to-day represents accurately the structural form of benzene. It has been established beyond doubt, by X-ray analysis of their crystals, that the aromatic substances graphite, hexamethylbenzene and hexachlorobenzene, and probably also benzene and naphthalene, possess a planar hexagonal structure of carbon atoms situated at a distance of 1.42 Å from each other, attached to which in co-planar orientation are the substituent groups. Similar results have been obtained by De Laslo by studying the diffraction of electrons by the vapours of aromatic substances, whereby it has been shown that the diameter of the benzene hexagon is 2.82 Å, and the carbon to carbon distance 1.41 Å.

The carbon to carbon distance within aromatic rings is, it will be noted, about 10% less than the corresponding atomic distances in the paraffin series (1.54 Å), slightly larger than that for a double bond (1.38 Å), and is, it hap-

¹ Bernal, Proc. Roy. Soc., 1924, A 106, 749; Lonsdale, Trans. Faraday Soc., 1925, 25, 352; Proc. Roy. Soc., 1929, A 123, 494; 1931, A 133, 536.

² Robertson, Proc. Roy. Soc., 1929, A 125, 542; 1933, A 142, 659.

⁸ Proc. Roy. Soc., 1934, A 146, 690.

pens, practically identical with the projection of the distance of separation of the atoms in the puckered ring structure of cyclohexane upon the median plane.

These measurements indicate that within an aromatic ring there must operate a powerful assemblage of interatomic forces which succeed in holding together the nuclei of the carbon atoms more closely than is possible within any aliphatic molecule. Further, the fact that the ultimate limit of aromatic ring condensation is found in the substance graphite (which is made up of hexagonal sheets of carbon atoms of honeycomb appearance, separated from each other by the large distance of 3.41 Å, across which the mineral can be sheared so exceedingly readily that it is an excellent lubricating substance) can be taken to indicate that the aromatic ring possesses some residual combining power, or "free affinity".

The formula I (a), however, represents the carbon atoms as being tervalent and not quadrivalent. To account for the fourth valency of the carbon atoms of the benzene ring is indeed the great problem of theoretical aromatic chemistry, which affects considerations not only of benzene but of all other ring systems of "aromatic character".

Kekulé himself suggested formulæ (II) (a and b) for benzenoid substances, but was not able to account for their lack of olefinic unsaturation. The absence of two isomeric ortho derivatives was however met, in 1872, by the suggestion that a continuous "oscillation of valency bonds" must be occurring between the alternative structures (a) and (b).

II (a)
$$X$$
 Y \rightleftharpoons Y (b)

Many syntheses of aromatic compounds from those of the aliphatic series indicate that benzene has a ring structure

¹ Compare Pauling, Brockway and Beach, J.A.C.S., 1935, 57, 2705.

^{*} Kekulé, Annalen, 1872, 162. 86.

containing three double bonds, though certain of them, as, for example, the synthesis of citrazinic acid ¹ or of orcinol, ² indicate a less symmetrical formula (III) which was indeed proposed by Dewar in 1867.



Perhaps the most significant of all these syntheses is that of Willstätter,³ who converted cyclohexanol successively into cyclohexene, cyclohexadiene and benzene. He showed that whereas distinct unsaturation resulted by introducing both one and two double bonds into the cyclohexane structure, yet when the third double bond was introduced there occurred a sudden change of properties with the appearance of the aromatic type.

This distinction in properties can be exemplified by comparison of the heats of combustion of the various compounds.

Heats of Combustion of Cyclic Hydrocarbons (in Kilogram calories per gram molecule)

·	-		•	_		Die.	erences.
n-Hexane .	•	• `	•		C_6H_{14}	989	
cyclo-Hexane	•	•			C_6H_{12}	937	52
cyclo-Hexene	•		•		C ₆ H ₁₀	892	45
cyclo-Hexadiene					$C_{\bullet}H_{\bullet}$	847	45
Benzene .		•			C_6H_6	78z	65

(b) Centric Formulæ.

After prolonged researches Baeyer b came to the conclusion that no definite stereochemical position could be assigned to the fourth valency bond of the carbon atoms

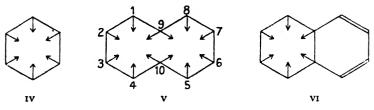
¹ Hofmann, Ber., 1884, 17, 2681; Ruhemann, ibid., 1887, 20, 799.

² Ingold, J. Chem. Soc, 1922, 121, 1143.

² Ber., 1912, 45, 1464. ⁴ Landolt-Bornstein, "Tabellen".

⁵ Ber., 1885, 18, 2278; Annalen, 1888, 245, 121; 1889, 251, 285.

in benzene, and hence suggested the "centric" formula (IV) with unlinked bonds.



This is indeed no more than a confession of ignorance with regard to the structure of benzene, satisfactory only in that it does represent the aromatic ring as having hexagonal symmetry. Centric formulæ are less satisfactory for condensed ring systems such as naphthalene or anthracene (V and VI), since they do not represent carbon atoms 9 and ro as being linked directly together (V), and give to no part of the molecule the six-membered ring structure which all reduction reactions prove to be present. Certain reactions, such as the formation of a diozonide only,1 the reactivity of the α -positions, and the facility of partial reduction to tetrahydronaphthalene have occasionally been pressed in support of the unsymmetrical formula (VI).2 This formula, however, is of very little intrinsic value, since it will be obvious that the moment any destruction of one of the aromatic rings of naphthalene has been initiated there remains the system of one benzene ring attached to a reactive unsaturated hydrocarbon chain.

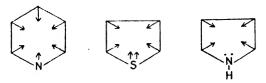
Over forty years ago it was pointed out by Bamberger ³ that although the unlinked bonds in the centric formula of Baeyer had no clear physical significance, yet any ring system of five or six atoms in which six bonds appeared to be dormant had distinctive aromatic character, no matter what the atoms constituting the ring. This extension of the centric formula from benzene to pyridine and thiophene is too well known to require comment, but application of

¹ Harries, Annalen, 1905, 343, 336.

² Willstätter and Waser, Ber., 1911, 44, 3430.

³ Annalen, 1893, 273, 373; 1890, 257, 47.

the same idea to pyrrole or to cyclopentadiene is much less clear.



The electronic implication of the centric formula of Baeyer and Bamberger, first indicated by Kaufmann in 1908, has been stressed by Robinson, who has stated that the characteristics of all benzenoid systems is to be sought in the hypothesis that six electrons are able to form a group which resists disruption. This group has been called the aromatic sextet.

In the case of pyrrole the sextet association requires both electrons of the unshared pair of the nitrogen atom, which can thus be no longer regarded as basic, although it can regain its normal condition by dissociation of proton, or by formation of, say, potassium pyrrole.

Cyclopentadiene and indene, too, are not truly aromatic systems, but they exhibit prototropy, and in the anions which result from the dissociation of proton the aromatic sextet again reappears. According to Robinson,³ who in this respect differs from W. Hückel 4 and others on the

¹ "Die Valenzlehre", 1908.
² J. Chem. Soc., 1925, 1604.

th Solvay Report, 1931; J. Soc. Dyers and Colourists. 1934, 69.
 Hückel, "Theoretischen Grundlagen der Organischen Chemie" (Leipzig, 1931), 1, chap. 9.

Continent, no structural significance should be associated with the existence of the sextet of electrons in benzenoid compounds, which should be regarded only as conjugated systems of a high degree of stability, owing their chemical inertness both to their stereochemical strainlessness and to association together of the six mobile electrons. The formula (VII) suggested by Robinson is therefore not to be regarded as a new structural formula for benzene, but merely as indicating stability in a cyclic system.



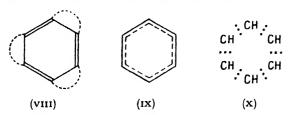
It has been pointed out 1 that the stability of beryllium acetylacetone, and, in general, all chelate co-ordination compounds, again can be ascribed to "aromatic character" and the possession of a sextet association of electrons.

(c) The Conjugation of Aromatic Rings.

From the foregoing sections it will be realized that no structural formula for benzene is in any way more satisfactory than the one originally proposed by Kekulé. This formula, however, as currently used at the present time, bears a certain specialized significance not explicitly stated in the original formulation. One of these implications is the recognition (due to Kekulé kimself) that within the ring system there can occur electromeric changes which increase

¹ Robinson and Armit, J. Chem. Soc., 1925, 1605.

the average symmetry of any aggregate of aromatic molecules. Others follow the recognition, by Thiele, that benzene is a completely conjugated system. The consequences of this important fact need elaboration, as they throw much light upon the nature of aromatic reactions.



By application of the hypothesis of the mutual saturation of free affinities in a conjugated system to the case of benzene. Thiele was able to account satisfactorily for its inertness towards the usual additive reagents for olefines, since the resulting formula (VIII) leaves the molecule with no residual free affinities, or points of potential attack. The formula (VIII) is, however, a reaction-formula, merely explicative of the Kekulé structure, and is not the equivalent of the structural formula (IX) in which the partial valencies are considered as real fractions of covalent bonds. The electronic formula (X), in which each of the bonds around the aromatic ring is represented as being constituted of three electrons, is the equivalent of (IX). Though suggested as long ago as 1914 by J. J. Thomson, 1 it is being used to-day by Kuhn, Hückel, and others on the Continent as the most modern representation of the structure of benzene.

From chemical considerations the partial valency formula (VIII) is more accurate than either (IX) or (X), because it represents benzene as differing from the olefines only in degree. It is equally satisfactory when applied to condensed aromatic systems.

The above statements are justified by the following important experimental facts:

- (i) In benzene the unsaturation is masked but is not non-existent. In sunlight benzene and chlorine combine to form C₆H₆Cl₆. With ozone the triozonide, C₆H₆(O₈)₈, is formed, and with diazoacetic ester addition occurs to a double bond.
- (ii) Benzene may be oxidized, by oxygen in presence of a vanadium pentoxide catalyst, to maleic acid. More frequently, however, derivatives oxidize in the 1:4 positions to yield quinonoid ring compounds, which again are fully conjugated.

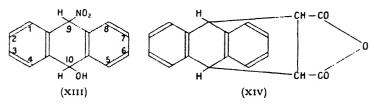
(iii) In naphthalenc the conjugation does not eliminate completely the "residual affinity" in the α -positions, which are indeed the initial points of attack in the molecule. Further, the only symmetrical formula containing two true aromatic rings, (XI), differentiates between the $\mathbf{1}:\mathbf{2}$ and the $\mathbf{2}:\mathbf{3}$ linkages. This is indeed shown up in differences of reactivity, as for example in the Skraup synthesis with

 β -naphthylamine, which yields only the naphthaquinoline (XII). This same difference is also shown by the fact that both β -naphthol and β -naphthylamine will couple in the 1- position, but not in the equally adjacent 2- position. In the heterocyclic bases quinoline and isoquinoline

414 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

these differences of reactivity are even more strongly marked.1

(iv) Anthracene cannot be given a structural formula which contains three aromatic rings: the meso positions, 9 and 10, most readily undergo additive reactions. Thus anthracene cannot be nitrated, but reacts with nitric acid to yield (XIII). Halogens, too, form addition- and not substitution-compounds. With maleic anhydride the Diels reaction occurs readily with the production of (XIV).² Only by oxidation can the molecule be stabilized, and then there results, in anthraquinone, a typical aromatic ketone containing two benzenoid rings. This olefinic character is even more strongly pronounced in polycyclic aromatic hydrocarbons, many of which behave as if they were free radicals.³



The conjugation of aromatic systems can be manifested not only by those reactions which are typical of addition to poly-enoid systems (p. 389) but also by numerous examples of ionotropic change. Thus p-nitrosophenol is tautomeric with p-benzoquinonemonoxime, and hydroxyazo dyes are indistinguishable from quinone-phenylhydrazones.

$$H = 0 + N = 0$$

In allied heterocyclic systems ionotropic change is even more frequent. Thus Decker 4 showed that γ -benzyl-

¹ Mills and Smith, J. Chem. Soc., 1922, 121, 2724.

² Diels and Alder, Annalen, 1931, 486, 191.

³ Clar, Ber., 1932, 65, 503, 1425.

⁴ Ber., 1905, 38, 3493.

pyridine (A) gives a yellow crystalline methiodide (B). When acted on with strong sodium hydroxide (over 20%) this is converted into an orange-coloured benzylidenemethylpyridane (C), which gradually becomes colourless after dilution to 5%, owing to the formation of the colourless hydroxide (D).

In the last action hydroxyl and hydrogen are added reversibly at the I:5 atoms of a conjugated chain, giving rise to a product which bears the same relationship to the original base as ammonium hydroxide does to ammonia.

Bases such as (C) are termed anhydronium bases. 1 and, as regards their chemical stability, appear to occupy a position intermediate between that of the fully aromatic true bases (e.g. D) and neutralized quinonoid systems. Though very diverse in chemical form, as can be seen from the examples given below, they all exhibit very similar chemical properties, diagnostic only of mobile conjugated systems.

chloride.

¹ Armit and Robinson, I. Chem. Soc., 1922, 121, 827.

416 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

9-Benzylidene-xanthane.

9-Benzyl-xanthylium chloride.

9-Benzylidene-thioxanthane.

9-Benzyl-thioxanthylium chloride.

It is unjustifiable, however, to ascribe the "aromatic" cha.acteristics of benzene solely to the conjugation of its

double bonds, for the analogous hydrocarbon, cyclooctatetrene (XV), C₂H₆, prepared by Willstätter in 1911,² is

¹ Perkin and Robinson, J. Chem. Soc., 1919, 115, 933.

² Ber., 1911, 44, 3428.

definitely olefinic in character, since it adds on bromine and reduces permanganate in the cold. From the stereochemical aspect, however, the two compounds C_0H_0 and C_0H_0 are quite distinct, for the latter must, if constituted of a planar structure of alternate double and single bonds, be a highly strained ring system, whereas Kekulé's formula for benzene is nearly strainless.

The pronounced effect of even a slight strain in a ring system has recently been demonstrated by Mills and Nixon, who point out that in the most stable state the angle α between a substituent group (or hydrogen) and a single bond of the carbon ring must be appreciably less than the angle β which it makes with the plane of the double bond on the other side (XVI). It follows that hydrindene

should be less strained, and therefore the more stable if of structure (XVIIa) rather than (XVIIb), whilst tetrahydronaphthalene should, on similar grounds, be more stable if of structure (XVIIIb) than (XVIIIa); the argument holding true even if it be admitted that the whole

¹ J. Chem. Soc., 1930, 2510.

molecule need not necessarily be uniplanar. In accordance with the consequences of this view it was found that 5-hydroxyhydrindene brominates and couples with diazocompounds in the 6- position, whereas β -(2)-tetrahydronaphthalene brominates and couples in the 1- position.

(d) Electronic Formulæ for Benzene.

The electronic formulation of Kekulé's structure for benzene (XIX) which is lacking in hexagonal symmetry, can be made more symmetrical either by following J. J. Thomson 1 and writing each of the carbon—carbon bonds with three electrons, or by following Robinson and indicating that six of the electrons might be differentiated from the rest so as to constitute a separate group (XX).

(X) and (XX) can be identified with the centric formulæ of Baeyer.

The substitution reactions of benzene, however, suggest that the molecule might have only three-fold and not six-fold symmetry, alternate atoms only having like properties as in any other conjugated chain. Thus dinitration or disulphonation of benzene yields preponderatingly the meta-disubstituted derivatives, rather than the 40:40:20 ratio of ortho, meta and para isomers that one would anticipate on statistical views for a symmetrical substance.

H. S. Fry,2 in 1912, considered this to be indicative of

¹ Phil. Mag., 1914, 27, 784.

² Cf. Collie, J. Chem. Soc., 1897, 71, 1013.

³ J.A.C.S., 1912, **34**, 664; 1914, **36**, 248, 262, 1035; 1915, **37**, 885. "The Electronic Conception of Valence and the Constitution of Benzene", 1921.

electrical polarization, and hence wrote the molecular structure of benzene as (XXI). At that date, however, G. N. Lewis had not published his classic paper discriminating between covalent and electrovalent linkages between atoms, and Fry did not therefore realize the error that he had made in superposing electrical charges on all the bonds, including those between carbon and hydrogen, whereas, according to Lewis, the integral charges of an electrovalence are alternatives to a covalence. The formula (XXI) nevertheless proved to be useful in explaining certain prominent experimental facts of aromatic substitution, e.g. the ortho — para substitution of phenol, aniline or chlorobenzene, and the meta substitution of nitrobenzene or benzoic acid (see p. 436), all of which are to-day adduced as evidence of the alternation of polarity within aromatic rings undergoing chemical change.

In 1923 Lowry ¹ brought the formulæ of Fry into accord with modern theories of valency by suggesting that benzene, which might well have the Kekulé, or any other, covalent formula when in the resting state, could be activated by the semi-polar dissociation of its double (or centric) bonds to yield (XXII) in which only the carbon atoms are charged, the positive and negative charges alternating round the ring.

¹ J. Chem. Soc., 1923, 123, 826.

This reaction-formula for fully-activated benzene accounts for all the chemical properties which the earlier structural formulæ were designed to express. Thus, the polarized formula (XXII) represents a molecule with practically no external electrical field, for the alternating positive and negative charges will give rise to a strictly localized system of Faraday tubes of force, like the lines of magnetic induction in an iron transformer ring. This is analogous to Thiele's partial valency formula with neutralized residual affinities. Again, neutralization of the electrical fields from positive and negative charges will occur not only around the ring, but, to a smaller extent, across it also, just as if there existed para or "bridge" bonds as in the formulæ of Claus and Dewar.

The conception that there is a weak electrical field in the vicinity of an aromatic ring can be used to account both for the weak adhesion between the successive planar layers of aromatic rings in crystalline graphite, and for the tendency of substances such as p-azoxyanisole to associate themselves into swarms of parallel molecules.

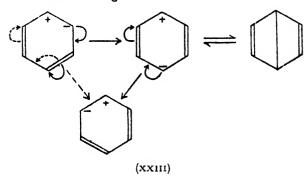
Extension of the Lowry formula for benzene to other condensed ring systems, such as naphthalene or anthracene, does not reveal any defects, for, even in the completely activated molecules the covalent six-membered rings all remain intact.



It is not necessary to assume that when a molecule such as naphthalene is activated every one of the double bonds is converted into a semi-polar bond. Only one ring might be activated, whilst the other might remain in the wholly covalent state. By supposing that activation may be localized in a complex system, one can account satisfactorily for the exceptional reactivity of the central rings of anthracene or phenanthrene.

Even in benzene one can suppose that the polar activa-

tion may be either partial or complete: controversy on any such minor point of theoretical speculation as this last is of no real consequence, but nevertheless it has divided organic chemists during the past decade. Whereas some, like Lowry, have preferred to consider that benzene should be either completely activated or completely inactive, others, such as Lapworth and Robinson, have supposed that the polar activation of an aromatic ring, which would necessarily develop from one double bond, does not spread automatically throughout the whole of the system, but yields a simple dipole, with two charges only, each of which, however, is prone to undergo electromeric migration to an alternate position. This migration probably would take place by electron movement around the ring, though it is also possible for electromeric changes to occur directly across the benzene ring.1



(e)-" Resonance" in Aromatic Rings.

The electronic formulæ for benzene, though quite satisfactory for the explanation of chemical reactions, do not give a correct representation of all physical properties.

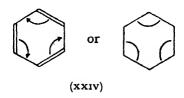
Aromatic compounds have very characteristic absorption spectra, situated usually in the near ultra-violet, which originate from rotations and vibrations occurring in the aromatic ring. Thus benzene gives a series of six sharp bands at regular intervals between 2000 and 2667 Å (com-

¹ Ingold, J. Chem. Soc., 1922, 121, 1133, 1143.

pare Fig. XV, p. 398), which can be correlated with the absorption band system of ethylene. The spectroscopic evidence, however, shows that all the links in benzene are the same, and thus indicates that the structure of a Kekulé formula (IIa) or (IIb) with both single and double bonds does not exist.

$$(11) \qquad \begin{array}{c} X \\ Y \\ (a) \end{array} \qquad \rightleftharpoons \qquad \begin{array}{c} X \\ Y \\ (b) \end{array}$$

Further, Pauling 3 has stated that the properties of benzene are not in accordance with those of a molecule possessing three electron links, as in Thomson's formula (X) (p. 412). Hence, one must conclude that benzene has a "resonance structure" (p. 396), in which the mutual perturbations of the electrons have resulted in the development of a stable symmetrical structure with a lesser energy content than any unperturbed structure with defined electronic positions. Thus the mesomeric system (cf. Fig.



XXIV) is to be regarded as more stable than any fixed structure.

The reason for the special stability of the benzene ring has been examined mathematically, and it has been shown

- ¹ Marsh, Phil. Mag., 1925, vi, 49, 1206.
- ² Bailey, Trans. Faraday. Soc., 1934, 30, 827. For an exhaustive analysis of spectral data regarding the structure of benzene see J. Chem Soc., 1936, pp. 912-987.
 - ³ Pauling, J.A.C.S., 1931, 53, 3225.
- ⁴ Hückel, Zeit. für Physik., 1931, 70, 204; 72, 310; 1932, 76, 628. Pauling, J. Chem. Physics, 1933, 1, 280, 362, 679. Penney, Proc. Roy. Soc., 1934, A 146, 223.

that the energy content of a benzenoid system possessing two alternative Kekulé structures will be lowered considerably if electronic interaction, or resonance, is presumed to occur. In the most stable structure it is not possible to associate all the electrons into pairs, for the six zp electrons have wave-functions which are closely coupled together. By considering all the electronic orbits it has also been shown that benzene is the most stable of all the cyclic hydrocarbons of formula C_nH_n .

The view that the aromatic ring is a structure stabilized by resonance has been supported by Ingold ¹ from chemical evidence.

(f) Summary.

In view of the many ways of depicting the same experimental facts, and the same underlying conceptions for their explanation, the following brief statement can be made with regard to the nature of aromatic rings, whether of benzene or of heterocyclic compounds.

- (i) All positions, and valency bonds, in benzene are exactly equivalent to one another, and the molecule probably owes its great stability to its high degree of symmetry, which is to a certain extent destroyed in heterocyclic rings and in condensed systems such as that of naphthalene.
- (ii) There is no reason why Kekulé's formula should not be used to-day for representing the structure of all aromatic molecules: it does represent in full the ways in which aromatic rings can react.
- (iii) Aromatic rings are fully conjugated systems, and owe their chemical stability, the reactivity of alternate atoms, and their tendency to undergo electromeric change to factors which characterize equally all other conjugated systems.
- (iv) Polar formulæ, either with + and signs (XXII), or with curved arrows to depict electron movements (XXIII), are intended to represent only specific reactions and not

¹ J. Chem. Soc., 1933, 1120.

424 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

general properties of aromatic compounds. (XXII) and (XXIII) are the exact equivalents of one another with regard to the chemical reactions they are used to describe.

AROMATIC COMPOUNDS—I

B. THEORIES OF AROMATIC SUBSTITUTION— AN HISTORICAL REVIEW.

(a) Early Studies.

Experimental investigation of aromatic substitution commenced at a very early date in the history of organic chemistry, and, naturally enough, the theoretical development of the subject followed attempts to classify together the substitution processes and the various aromatic compounds that could be produced through them. Until the correct structural relationships between the *ortho*, *meta* and *para* disubstitution derivatives of benzene had been established no systematic study of aromatic substitution was possible, but, very shortly after this work had been completed, attempts at formulation of orientation rules were made by Hübner in 1875, and by Noelting in the following year.

It was clearly recognized by these investigators (i) that substitution of benzene derivatives in the *ortho* (I:2) and *para* (I:4) positions occurred together, without simultaneous substitution in the *meta* (I:3) position (which alone was distinctive of the substitution of another class of compounds), and, (ii) that the mode of substitution was dependent essentially upon the substituent group already present, rather than upon the nature of the group being introduced.

Noelting, indeed, went further, and connected the orienting effect of a substituent group with its chemical character, by stating that *meta*-directing groups were of a "strongly acid" character. When one considers that, at that time, ideas of acidity and basicity were still expressed in the vague terminology of the already discredited electrochemical theories of Berzelius, one realizes that no more

explicit exposition of orienting effects could then have been given. Indeed, modern theories of aromatic substitution are to a great part concerned with a fuller electronic interpretation of this same connection between acidic or basic character of substituent groups and directive action.

(b) Significance of Valency.

A second aspect of the subject was revealed in 1887, when H. E. Armstrong ¹ pointed out that, in general, radicals with monovalent linkages only, such as —CH₃ or —Cl, were ortho — para-directing, whereas those groups containing double or triple bonds, such as —CO·OH or —C=N, were meta-orienting. Unsaturation therefore appeared to be a characteristic of a meta-directing group. This correlation was again stressed by Vorländer ² in 1902. Later experimental investigation has shown that exceptions to the rule of Armstrong are not infrequent; azobenzene, C₆H₅—N=N=N—C₆H₅, and cinnamic acid, C₆H₅—CH=CH—CO·OH, which are both ortho — para substituted, affording typical examples.

(c) The Crum Brown—Gibson Rule.

The formulation of a general "Substitution Rule" by Crum Brown and Gibson in 1892 may be said to close the period in which experimental data were collected together merely for the derivation of empirical summaries applicable for the prediction of the structures of the products of substitution reactions. These authors pointed out that whenever an aromatic compound C₆H₅X contains a group X whose hydride is easily oxidizable by a one-stage process to HOX then C₆H₅X gives meta-substitution derivatives, whereas if HX is not directly oxidizable to HOX then X is an ortho — para-directive grouping.

The above rule is a very convenient and accurate empiricism, for, in fact, the very few criticisms that have been made against it 4 have arisen only through the difficulty

¹ J. Chem. Soc., 1887, 51, 258. ² Annalen, 1902, 320, 122.

³ J. Chem. Soc., 1892, 61, 367.

⁴ Cf. Bone and Drugman, J. Chem. Soc., 1906, 89, 679.

of certifying whether or not a compound HX is "oxidizable by a one-stage process" to HOX. Actually the rule, as formulated above, is not of much theoretical use, for it does not focus attention on the true functional characteristic of a substituent which renders it either ortho — para or meta directive.

Mention has already been made of Vorländer's effort ¹ to explain the correlation by reviving the earlier rule of Armstrong and connecting the ease of oxidation of a compound HX with the saturated or unsaturated character of X. This, however, failed on account of the limitations of Armstrong's original conception.

Nevertheless, if the generalization of Crum Brown and Gibson is expressed in the following less definite manner (which undoubtedly was implicit in the original formulation of the substitution rule) stating that

"if a group X is more stable in its compounds with hydrogen than with hydroxyl . . . cf. $HCl \cdot HOCl : H \cdot NH_2 \cdot HO \cdot NH_3 \cdot H \cdot OH \cdot \cdot \cdot HO \cdot OH : . . . then it is an ortho — para-directing substituent, whereas if the reverse is the case . . . cf. <math>HO \cdot NO_2 \cdot H \cdot NO_2 : HO \cdot CO_2H \cdot H \cdot CO_2H \cdot . . . then X is a meta-directive group ",$

it is seen that the underlying factor determining the mode of action of a substituent in an aromatic compound is its polar electrical character. Through the realization of this fact by H. S. Fry, Vorländer, Lapworth, and numerous later workers, it has been possible to incorporate all these empirical classifications into the general scheme of modern theoretical organic chemistry.

The following generalized "substitution rule" was put forward by Hammick and Illingworth in 1930:

"If in the benzene derivative C_6H_5 -XY, Y is in a higher group of the periodic table than X, or if, being in the same

¹ Annalen, 1902, **320**, 122. ² J. Chem. Soc., 1930, 2358.

428 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

group, Y is of lower atomic weight than X, then a second atom or group of atoms that enters the nucleus will do so in the *meta* position to the group XY. In all other cases, including that in which the group XY is a single atom, a second entering atom or group goes to the *ortho* and *para* positions. The effect of ionic charges on XY is given by the statement that a positive charge directs *meta*, and a negative charge directs *ortho—para*."

However, Masson 1 has pointed out that

"The rule of Hammick and Illingworth, however, is valid only so far as it may correctly imply which elements, X, in the Periodic Classification are positive to which other elements Y; it fails when, as in the case of $-XY = -IO_z$, X can be said to gain more of positive character from being in a later Period than of negative character from being in a later Group."

(d) Direct and Indirect Substitution.

Many investigators had, before 1900, noticed that the introduction of a substituent into an aromatic ring was often the result of a series of reactions involving the molecular rearrangement of an intermediate substance, and had suggested tentatively that there might be a genetic connection between aromatic substitution and isomeric change. Thus aromatic substitution might be an indirect process in which the directing substituent group was primarily concerned. Blanksma² brought these views into prominence in 1902; he suggested that the substitution of a omatic compounds containing the groups —OH, —NH₂, or those derived from them, was an indirect process, in which the substituting reagent first attacked the "sidechain" substituent already present, giving an unstable product which readily underwent isomeric change to a ringsubstituted product. He also pointed out that the substitution of aromatic substances containing these groupings, occurred with very much greater ease than did all other substitution reactions, and gave exclusively ortho -- parasubstituted products. A later extension of this theory to all rapid ortho - para-substitution reactions has not been

¹ Masson, Race and Pounder, J. Chem. Soc., 1935, 1673.

² Rec. trav. chim., 1902, 21, 281; 1904, 23, 202.

justified by experimental facts.¹ In addition, it has been shown that the molecular rearrangement of aromatic compounds are by no means simple processes (pp. 474–483). Nevertheless, the hypothesis of indirect substitution is still valuable in the limited sphere for which it was originally proposed.²

Further, in the development of our knowledge of the nature of aromatic substitution processes, Blanksma's suggestion has been valuable because it drew attention to the necessity of taking into account the way in which a substitution reaction occurs as well as the final results which it gives. Also, it indicated that the mechanisms of the processes of isomeric change might be of significance in connection with the elucidation of the reactivity of aromatic ring compounds.

(e) The Conception of Alternating Affinity.

In 1902 Flürscheim, by adapting Werner's theories of valency, including in particular its utilization of Faraday's conceptions of the distribution of tubes of force in space, suggested that an alternately strong and weak distribution of "chemical affinity" around an aromatic ring could be induced by the "affinity demand" of substituent groups. This unequal distribution of chemical affinity would allow aromatic substitution to occur either at the ortho and para positions or else at the meta position according to the nature of the substituent group already present in the aromatic ring.

Groups like chlorine or methyl were supposed to make a strong affinity demand on the carbon atom (1) to which they were united and to leave, in consequence, "free affinity" at carbon atoms 2, 4, and 6, facilitating substitution at these points. Groups like —NO₂, in contrast, were supposed to exert only a weak affinity demand on carbon

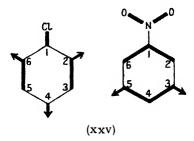
¹ Cf. Hollemann, "Die Direkte Einführung . . .", 1910, p. 197.

² Cf. Francis, J.A.C.S., 1925, 47, 2230; 1926, 48, 1637; De

²Cf. Francis, J.A.C.S., 1925, 47, 2230; 1926, 48, 1637; De Lange, Rec. trav. chim., 1926, 45, 19.

³ Flürscheim, J. prakt. Chem., 1902, 66, 321; 1905, 71, 497; Ber., 1906, 39, 2015.

atom *I*, and to create points of reactivity at carbon atoms 3 and 5. The actual distribution of chemical affinity within any aromatic molecule was indicated by means of formulæ such as those shown below (Fig. XXV), in which thick and thin lines indicated bonds utilizing large and small quantities of chemical affinity respectively, whilst arrows attached to bonds not linking two atoms indicated "residual", or "free", affinity, available at carbon atoms which would in consequence be expected to be chemically reactive.



It was supposed by Flürscheim that the whole of the chemical affinity of any one atom had to amount to a constant total, just as the number of Faraday tubes of electrical force proceeding from any electrical charge had to amount to a fixed total. At the same time it was emphasized that the chemical affinity of any one chemical bond need not be constant, unless it was a bond linking a univalent element, such as hydrogen, to any other atom. Owing both to Flürscheim's unorthodox views of the nature of valency bonds, and to his rather confusing diagrammatic representations, his speculations met with much criticism.1 particularly from later workers accustomed to the electronic theory of valency which does not allow of the postulation of covalent bonds of variable character. Nevertheless, the underlying conception can be modified, as its originator has shown,2 so as to give a valuable insight

¹ Cf. Obermiller, J. Prakt. Chem., 1908, 77, 78; Hollemann, ibid., 1906, 74, 157; "Die Direkte Einführung . . .", pp. 211-12 (1910); Tschitschibabin, J. Prakt. Chem., 1912, 86, 397.

² Chemistry and Industry, 1925, 44, 246; J. Chem. Soc., 1926, 1562.

into the mode of action of electrical forces within molecules.

The idea that alternations in degree of chemical reactivity can be transmitted inductively from a substituent group situated in a distant part of a molecule is of fundamental importance.

(f) Hollemann's Theories.

A new stage in the development of the subject of aromatic substitution was marked by the publication, in 1910, of a book entitled "Die Direkte Einführung von Substituenten in den Benzolkern", which contained a systematic description of all the then known facts about aromatic substitution reactions. In it A. F. Hollemann emphasized the value of making exact quantitative measurements of the percentage yields of all the isomeric products of aromatic substitution reactions.

From classification of the substitution reactions of disubstituted benzene derivatives, C₆H₄XY, Hollemann was able to show which had the stronger orienting effect of any two groups. He indicated, as follows, the relative magnitudes of the directing powers of substituents:

Ortho - para-Directing Groups.

The above list may be used to-day, with amplification by insertion of other substituents, to predict the nature of the product to be obtained on substitution of practically any aromatic compound. This may be illustrated by the

following examples:

(i) If p-toluidine is brominated, the bromine enters the ring in the positions in the ortho relationship to the powerful directing group —NH₂ in preference to positions in the ortho relationship to the weakly directing group —CH₂.

(ii) If m-nitro-toluene is further nitrated, the nitro-group enters positions in the ortho and para relationship to the —CH₃ group and not the vacant position in the meta relationship to the —NO₂ group, as this has no appreciable directive influence.

Any table such as the above is a certain guide to the prediction of the results of substitution processes, because it is compiled from the ever-growing basis of experimental facts, and is not dependent upon theoretical speculation. Later investigations have shown, however, that, under certain circumstances, the relative orders of certain substituent groups in Hollemann's table can be inverted for some, though not for all, reactions.¹

It may be noted that Hollemann classified substituents according to their "directive power", which he measured by the relative speeds of substitution under standard conditions. Most frequently, the reaction studied was the nitration at o° C. by fuming nitric acid. Hollemann himself stressed the important fact that whereas ortho — paraorienting substituents tended to increase the rate of aromatic substitution, as compared with benzene itself, the metaorienting groups greatly decreased the rate of reaction. For example, whereas benzene has to be nitrated by the use of a mixture of concentrated (ca. 60%) nitric and sulphuric acids, phenol reacts with dilute nitric acid with great vig ur, whilst benzoic acid can only be nitrated by the use of fuming nitric acid, or solid potassium nitrate, in the presence of a large excess of hot concentrated sulphuric acid.

Substitution can, theoretically, occur at every point in an aromatic substance, and it is even possible to replace substituent groups other than hydrogen,² but, except in the case of symmetrical substances, substitution actually proceeds at rates which differ very widely in different positions. The positions taken up by the substituting groups depend

¹ Cf. Ingold and Vass, J. Chem. Soc., 1928, 417.

² Orton, J. Chem. Soc., 1901, 79, 822; 1907, 91, 1543; Francis, J.A.C.S., 1924, 46, 2498; 1925, 47, 2211; De Lange, Rec. trav. chim., 1926, 45, 19.

upon the velocity with which substitution can take place—the most reactive point being chosen preferentially but not exclusively. Hence substituent groups already present within an aromatic molecule must act by modifying the velocity of substitution at distant points.

Obermiller 1 pointed out that one can consider an initial weakening of the carbon to hydrogen covalent bond, by induced influences, to be an essential stage in any substitution reaction. Alternatively Hollemann, following H. E. Armstrong, suggested that aromatic substitution involved the initial addition of two groups to an unsaturated linkage of the aromatic ring, followed by the elimination of a simple inorganic molecule such as water or hydrogen bromide:

According to this addition hypothesis, ortho — para-orienting groups would be those which favoured addition at positions conjugated to them (I:2-addition leading to ortho substitution and I:4-addition to para substitution) whilst the meta-substituting groups might be those which inhibited reaction at these points. In accordance with this idea it was pointed out that when meta substitution occurred the principal by-product was the ortho isomer, as would be expected with an addition process.

There exists a considerable amount of experimental evidence to support the statement that some, though not all, of the typical aromatic substitution reactions occur in stages, the first of which is an addition process. In particular, nitration may occur by a mechanism involving the initial addition of the groups —NO₂ and —OH to an unsaturated system. Many olefines can be nitrated if the requisite conditions, in which oxidation is avoided, are

found, and occasionally the intermediate hydroxy-nitro addition-compound can be isolated. E.g. ¹

$$(C_{\mathfrak{g}}H_{\mathfrak{s}})_{\mathfrak{g}}C == CH_{\mathfrak{g}} + HO - NO_{\mathfrak{g}} \longrightarrow (C_{\mathfrak{g}}H_{\mathfrak{s}})_{\mathfrak{g}}C - CH_{\mathfrak{g}} \longrightarrow (C_{\mathfrak{g}}H_{\mathfrak{s}})_{\mathfrak{g}}C == CH \cdot NO_{\mathfrak{g}}$$

The action of nitric acid on anthracene is undoubtedly of similar type:

In this connection, it is significant that a small quantity of picric acid is always obtained by vigorous nitration of benzene, and if mercuric nitrate be added as catalyst, can be produced in up to 50% yield, the process being of commercial value in explosive manufacture. The production of a nitrated phenol from the hydrocarbon can be explained simply by postulating addition of hydroxyl groups to the nucleus.

The Friedel-Crafts reaction (p. 231) and the coupling reactions of phenols and amines also seem to occur by processes involving primary addition. Thus phenol ethers will combine with diazonium salts, giving ultimate products in which free hydroxy-azo dyes are present:

OCH₃

$$+ Ph \cdot N_2 \cdot OH \rightarrow H OCH_3$$

$$+ Ph \cdot N_2 \cdot OH \rightarrow H N_2 Ph$$

$$+ H_2 O$$

$$+ N_2 Ph$$

¹ Wieland and Rahn, Ber., 1921, 54, 1770; Wieland and Sakellarios, ibid., 1920, 53, 201.

In bromination, too, phenols act as tautomeric ketoenolic substances capable of forming addition-compounds. For example, β -naphthol will brominate instantly in the adjacent α -position, but if this is already substituted, a bromo ketone is formed which, in presence of hydrobromic acid, can isomerize to a 6-bromo- β -naphthol.¹

Here, again, anthracene gives a similar reaction with bromine.

There is not, however, sufficient experimental evidence to warrant the assertion that intermediate addition-compounds are always formed in aromatic substitution. It seems far more probable that many substitution reactions involve the direct attack of a reagent at a single carbon atom, with fission of a C—H linkage.

(g) Electronic Theories.

Even before G. N. Lewis had drawn the distinction between electrovalent and covalent bonds, H. S. Fry³ attempted to explain the alternation in chemical character of the atoms of any substituted benzene derivative by suggesting that positive and negative charges actually resided upon the atoms constituting the aromatic molecule. He

¹ Fries and Engel, Annalen, 1924, 439, 232.

² Barnett and Cook, J. Chem. Soc., 1924, 125, 1084.

⁸ Zeit. phys. Chem., 1911, 76, 385, 396, 591; J.A.C.S., 1919, 34, 664; 1914, 36, 248, 1035; "The Electronic Conception of Valence and the Constitution of Benzene", 1921.

put forward the structure (XXI) for benzene, and, by considering substitution reactions to be merely ionic interchanges, it followed directly that reaction would occur at alternate positions according to the electrical sign of the hydrogen atom that could be displaced. Fry realized that the hydrogen atoms of most electropositive character were those which could be substituted most easily, since they

were liberated in the form of HOH or HBr.

He pointed out that the substituent groups already present induced, in virtue of their inherent electrochemical characters, the occurrence of these specific reactive electropositive hydrogen atoms at certain orientation positions only. Thus chlorine, being essentially an electronegative element, induced the occurrence of positive hydrogens at the ortho and para positions, whilst the nitro-group induced the occurrence of positive hydrogen atoms at the meta positions.

Subsequent to the recognition of the non-polar character of covalent bonds, the theory of Fry was, in its original form, no longer tenable, but certain aspects of it are still valuable, since it is now universally accepted that the carbon atoms of the benzene ring must acquire distinctive polar characteristics at the moment of chemical change, if even at no other time.

Vorländer 1 in 1919 combined the theories of Fry and Flürscheim in a very simple form to give a substitution rule entirely consistent with experimental fact. He stated that aromatic compounds C₄H₄X containing an atom of

¹ Ber., 1919, **52**, 263; 1925, **58**, 1893.

electropositive character attached directly to the aromatic ring are all inert and meta substituting, whilst those with an electronegative atom adjacent to the aromatic ring are reactive and ortho — para substituting (cf. p. 427). The rule can be given still more precisely as follows: "Positive poles, or the positive ends of permanent dipoles, are, when attached directly to aromatic rings, meta-directing substituents, whereas negative poles, or the negative ends of dipoles, are ortho — para directing."

Vorländer laid particular stress on the important fact that it is the atom linked directly to the benzene ring which has the most influence upon its mode of substitution, pointing out that one could vary the side-chain of an aromatic compound without changing its type of nuclear substitution. Thus the ortho - para-directive effect of a hydrocarbon linkage, as in toluene, is preserved in tertiary butyl benzene, C₆H₅·C(CH₃)₃, benzyl chloride, C₆H₅—CH₂Cl, diphenyl, C₆H₅—C₆H₅, and all its derivatives, phenylacid, C₆H₅—CH₂·COOH, and cinnamic acid. acetic C_aH_s—CH—CH-COOH. Very drastic substitution of the side-chain indeed, such as the introduction of the C--O group, or of a -NO₂ group, in the α-position (phenyl nitromethane is on the whole meta substituting), is needed to alter the sign of the resultant dipole of the C₆H₆—C linkage.

Although Vorländer's generalizations will predict exactly the results of an aromatic substitution process, they do not explain its real mechanism. The credit for the consistent incorporation of our modern electrical conceptions of the nature and reactivity of chemical molecules into theories of aromatic substitution must be given to English investigators.

In 1920 Lapworth ¹ applied his earlier theories of polarity and chemical change (pp. 380-392) to the case of aromatic substitution, and explained the production of reactivity at particular points in a benzene ring, by substituent groups of polar character, by a theory of *induced* electrical polarity.

¹ Lapworth, Mem. Manchester Phil. Soc., 1920, **64**, iii, 2; J. Chem. Soc., 1922, **121**, 416.

According to Lapworth there could be transmitted within any aromatic ring, at the moment of reaction though not necessarily in the compound before activation (cf. p. 386), an electrical polarization like in sign on alternate atoms only, just as there could be made to occur in conjugated tautomeric systems like points of polar reactivity on alternate atoms (cf. Lapworth's rule, p. 380).

The presence of a key atom of definite polar character was necessary in a "directing group" before this transmission of electrical polarity occurred preferentially in any one specific way, and, the more pronounced the polar character of the key atom, the more would substitution be restricted to the one type. Thus, the very weakly electronegative carbon atom, or the weakly electropositive hydrogen atom, of the methyl group in toluene favoured ortho—ara substitution, but meta substitution still occurred to as much as 4% of the whole, whilst, in contrast, when a powerful key atom, such as the oxygen atom in phenol, was present, substitution became restricted exclusively to the one type.

Lapworth's views were amplified in 1922 by Kermack and Robinson, who gave a much clearer pictorial representation of how electrical polarization could be propagated in a conjugated chain. Subsequent developments of the theory by Robinson, Ingold, and others, have been very extensive, but no completely fresh ideas have been introduced, although the original scheme has been modified in several particular details.

The modern viewpoint with regard to aromatic substitution does represent a general consensus of opinion. It is discussed in some detail in the following sections of this chapter.

¹ Kermack and Robinson, J. Chem. Soc., 1922, 121, 427.

² The following are some of the more important of the recent papers dealing with theories of aromatic substitution: Allan, Oxford, Robinson and Smith, J. Chem. Soc., 1926, 401; C. K. Ingold, and E. H. Ingold, J. Chem. Soc., 1926, 1310; Ingold, Chemical Society Annual Reports, 1926–28; Robinson, Solvay Report, 1931; Lectures to Institute of Chemistry, 1932; J. Society of Dyers and Colourists, 1934, 65.

AROMATIC COMPOUNDS—I

C. THE MECHANISM OF SUBSTITUTION PROCESSES.

(a) Normal Substitution of Nuclear Hydrogen.

From the foregoing historical review of theories of aromatic substitution there emerge the following general facts concerning the actual substitution processes:

(i) In normal substitution the entrant group displaces hydrogen, which is either liberated as proton, H⁺, or else becomes part of a substance, such as water, appreciably more acidic than the aromatic hydrocarbon. E.g.

The entrant group is thus a *kationoid* group, and must react with an anionoid centre in the aromatic nucleus.

- (ii) The first stage of an aromatic substitution reaction resembles that of addition to an olefine (Chapter VIII), where again an anionoid carbon atom is the initial point of attack. In most cases no "addition-compounds" can be isolated, and, whenever this is the case, it is immaterial whether the subsequent stages of aromatic substitution are confined to "attack at one point"—i.e. to the elimination of proton—or whether they include complete addition of the second radical of the substituting reagent, followed immediately by a cyclic process of re-elimination together with a proton.
- (iii) The process by which the nuclear carbon atom at the point of substitution becomes activated to an anionoid centre, through the directing agency of the substituent group already present in the aromatic ring, may be identical with that by which electromeric changes occur in conjugated systems.

440 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

(iv) Substituent groups, of definite polar type, can affect the velocity of reaction throughout the whole of an aromatic nucleus.

These facts are consistent with the theories of polar reactivity that have been developed in earlier chapters of this book, and hence there is no necessity to regard aromatic substitution as a special process. Thus, the replacement of nuclear hydrogen by another kationoid group can be represented clearly by the following diagram, due to C. K. and E. H. Ingold, which suggests that the actual reaction is largely completed, though not initiated, by the mutual influence of two dipolar molecules:

The above representation of aromatic substitution as a mutual interaction of two dipoles needs supplementation, however, by some indication of why particular carbon atoms in an aromatic substance can validly be regarded as potential anionoid centres of reactivity, since with any carbon—hydrogen bond there is associated only a negligibly small dipole moment. Two hypotheses may be used:

(i) A carbon atom may be situated in an electrostatic field of negative potential that can be ascribed to the "general polarity" (Chapter XI) of a substituent group. E.g.

$$O \longrightarrow H$$
: $CH_3 \longrightarrow H$

$$-F(+I)$$
Phenoxide anion. Toluene.

(ii) A double bond in the aromatic ring may become activated, spontaneously, or with a local catalyst as in the

¹ J. Chem. Soc., 1926, 1313; Chemical Society Annual Reports, 1926, 23, p. 135.

case of ethylene (p. 172) or, again, the negative pole may reach a definite carbon atom by an electromeric change, as in the case of butadiene (p. 372). E.g.

HC CH HC CH HC CH

HC CH
$$\rightarrow$$
 HC \rightarrow CH

HC CH

HC CH

HC CH

HC CH

(a) Simple activation.

(b) After electromeric change.

The methods (i) and (iib) by which an aromatic ring could be polarized were termed Inductive (I) and Tautomeric (T) processes respectively by C. K. Ingold in 1926. The latter activation process frequently originates at a substituent group, when the aromatic ring becomes activated only after an extended electromeric change. E.g.

$$H_2 \stackrel{CH}{\longrightarrow} CH \stackrel{CH}{\longrightarrow} CH \stackrel{CH}{\longrightarrow} CH \stackrel{CH}{\longrightarrow} CH$$

Nuclear carbon inert.

Nuclear (p) carbon reactive.

The speed of an aromatic substitution initiated by an electromeric change, such as that illustrated above, will be proportional to the chance that, at the particular instants when the selected carbon atom is an active anion there is in close proximity a reagent substituent kation, and of course proportional also to the frequency with which these moments of polar activation occur.

It is generally agreed to-day that the electromeric activation of aromatic substances is usually initiated at the preexisting centres of electrostatic polarization in the molecule. The directive action of a polar substituent group is explained more fully in the following sections.

¹ Ingold, Chemical Society Annual Reports, 1926, p. 140; cf., however, J. Chem. Soc., 1932, 1445, 2631.

(b) Substitution of Amines and Phenols . . . in general, Hetero-enoid Systems.

Phenol affords a typical example of a hetero-enoid system (p. 390), for it contains an atom of variable valency (oxygen) linked directly to the conjugated hydrocarbon ring.

Hence it can be ortho or para substituted by the following cycle of processes.

(i) Activation and Electromeric Change.

The oxygen atom in phenol possesses lone pairs of electrons, which on activation become mobile, when electromeric change is apt to occur.

$$H = \stackrel{\circ}{0} =$$

This can lead (a) to the production of a negative pole at the *ortho* position, or (b) to the production of a negative pole at the *para* position. The activated aromatic entities (a) and (b) are, it must be noted, structures of quinonoid type.

(ii) Interaction with a Kationoid Reagent.

A kationoid reagent, which need not necessarily be already ionized, can combine with the activated anionoid carbon atom.

$$H \longrightarrow C \longrightarrow C \longrightarrow C \longrightarrow C \longrightarrow Br + \overline{Br}$$

$$H \longrightarrow C \longrightarrow C \longrightarrow C \longrightarrow Br + \overline{Br}$$

(iii) Elimination of Proton and Electromeric Change.

The elimination of proton from the product of reaction (ii) yields a quinonoid entity which can become stabilized by a sequence of electromeric changes which are the exact reverse of those of process (i). The final product is the electrically-neutral substituted phenol.

As Robinson has pointed out, the electromeric changes in the nucleus, and the involved external group, have, from the viewpoint of the substitution reaction as a whole, an oscillatory character. It has already been suggested (p. 316) that polar molecules may complete the electrical circuit in the solution, and reduce greatly the energy needed for the initial activation.

Electron movement from substituent to nucleus, as described above, can only occur when the substituent group possesses unshared electron pairs. In addition to phenols and amines, the halogens can conceivably react in this way. Quinonoid structures, such as

$$: \dot{C}_1 = C \\ C = C \\ C \\ H$$

are quite unknown, although it is quite possible for an atom of a halogen element to lose an electron pair (p. 167). The activated structures indicated in the above scheme only indicate momentary *phases* in the whole reaction, and not ionized molecules having any appreciable free life.

¹ Chemical News, 1926, **133**, 3455; Institute of Chemistry Lectures, 1932; J. Society of Dyers and Colourists, 1934, 69.

444 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

The speed of substitution of the hetero-enoid aromatic systems can be correlated with the mobility of the unshared electron pairs in the substituent group. Hence, in accordance with the schematic review of basicity given in Chapter IX one can place the efficiency of the ortho — para-directive groupings in the following order:

$$\bar{RN} > \bar{O} > R_{\bullet}N > RO > I > Cl.$$

Among derivatives of amines the order will follow the order of basicity N(Alkyl)₂>NH₂>NHAcyl, and similarly amongst phenol derivatives the order will be

O>OH>OAcyl. Amongst the halogen elements the order is I>Br>Cl>F.³ Lapworth and Robinson have suggested that para substitutions may be regarded as due to frequent activations at a low energy level, and ortho substitutions as due to a less frequent and more energetic activation.

Of the numerous experiments that have been carried out to compare rates of aromatic substitution, the measurements made at Bangor University with phenols and their derivatives deserve special mention.

Free phenols are chlorinated smoothly by hypochlorous acid, and in acid or alkaline solution the reaction velocity, which can be measured accurately by iodometric titration, is in direct proportion to the hydroxyl ion concentration of the mixture, indicating that the rapid aromatic substitution takes place by attack of phenoxide anions by hypochlorous acid molecules.⁵ For different phenols it appears that the product

 $Velocity\ of\ Chlorination\ imes\ Ionization\ Constant$ is roughly constant, indicating that the anions derived

¹ Note.—Hollemann's sequence is OH > NH₂ > Cl (p. 431), but in practice he compared the groups NHAcyl and OAcyl.

² Cf. Lapworth, J. Chem. Soc., 1922, 121, 416; Allan, Oxford, Robinson and Smith, ibid., 1926, 403; Ingold and Shaw, ibid., 1927, 2922; Robinson, J. Society of Dyers and Colourists, 1934, 73.

³ Cf. pp. 257-8.

⁴ Mem. Manchester Phil. Soc., 1927-8, 72, 43.

⁵ Soper and Smith, J. Chem. Soc., 1926, 1582.

from the more acidic (i.e. less basic) phenols are the less reactive.

TABLE I

Velocity of Chlorination of Phenols 1

	Chlorination	Ionization	Product
Phenol.	Velocity (V).	Constant (K).	$\mathbf{K} \times \mathbf{V}$.
Phenol	3.3×10^{5}	$1 \cdot 1 \times 10^{-10}$	3.63 × 10 ⁻⁶
o-Cresol .	1.25×10^{5}	o·63 × 10 ^{−10}	7.88×10^{-8}
m-Cresol .	4.4×10^{5}	0.98×10^{-10}	42.00 × 10-5
p-Cresol .	9.0 × 10 ⁵	0.67×10^{-10}	6·0 × 10 ⁻⁸
o-Chlorophenol	5.9 × 104	7.7×10^{-10}	4.55 × 10-6
p-Chlorophenol	4.90 × 104	4·I × 10-10	2.0 × 10-8
o-Nitrophenol	$1.01 \times 10_{8}$	7·5 × 10−•	7·58 × 10-1
m-Nitrophenol	5.89×10^3	1.0 × 10-	5.89 × 10-5
p-Nitrophenol	6.84 × 10°	9·6 × 10~*	6.57×10^{-6}

Chlorine is even more reactive than hypochlorous acid, and reacts with free phenols at an unmeasurably high speed. From experiments with phenol ethers it appears again that the velocity of halogenation, which occurs exclusively under the directive influence of the alkyloxy group, is controlled by constitutional factors affecting the basicity of the oxygen atom, quite independently of any other factors (e.g. field effects) originating from other parts of the benzene nucleus.² This may be seen from the following tables, in which are compared (II) the rate of chlorination of compounds R·O·C₆H₄·X, compared to CH₃·O·C₆H₄·X as 100, and (III) rates of chlorination of R·O·C₆H₄·X compared to p-R·O·C₆H₄·COOH as 100.

TABLE II

R =	Me.	Et.	Pr(*).	Pr(iso).	Butyl(#).	Benzyl.	p-NO _s - Benzyi
x							
p-COOH	100	198	215	444	221	70	
p-NO.	100	200	221				
p-NO, p-Cl	100	199	225	439	222	67	14.2
o-Cl	100	199					14.4
p-Cl	R = ;	p-Chlore	benzyl,	39.4:	m-Nitro	benzyl,	17.2

¹ Soper and Smith, J. Chem. Soc., 1926, 1582.

² Bradfield and Jones, J. Chem. Soc., 1928, 1006.

TABLE III

	X = COOH.	C1.	NO.
R = Methyl	100	276	0.674
Ethyl	100	278	0.682
n-Propyl	100	288	0.691
iso-Propyl	100	272	
n-Butyl	100	283	
Benzyl	100	269	

Comparison of the columns of Table II shows the change of efficiency of the electromeric activation of chlorination by the group $R \cdot O$ — to be a constant for any group $R \cdot$, independent of other substituent groups, -X, situated elsewhere in the molecule. Table (III) reveals the equally important fact that the electrostatic field (-I or + F) of a group -X, not controlling the mode of substitution, does affect the reaction velocity of the molecule by a constant, and yet independent, amount. The independence of the two factors controlling the rate of aromatic substitution has been confirmed by a wide range of quantitative experiments.

(c) Salt Formation in Arematic Substitution.

The electromeric change that can be initiated by the lone pair of electrons of the nitrogen atom of an amine cannot occur in an ammonium salt. Consequently it is not surprising that aromatic ammonium salts are, in sharp contrast to the amines, inert *meta*-substituting compounds. Thus both aniline and dimethylaniline react instantly with bromine in aqueous solutions to give the tribromo derivatives, whilst the trimethyl-phenyl ammonium salts, $C_{\bullet}H_{\bullet}$: $N(CH_{3})_{3}X$ are inert. They can be nitrated with

C₆H₅·N(CH₃)₃X are inert. They can be nitrated with difficulty, using concentrated nitric and sulphuric acids, giving *meta*-nitro derivatives. In concentrated sulphuric acid solution, aniline itself can be nitrated to *meta*-nitraniline.²

¹ Bradfield and Jones, *J. Chem. Soc.*, 1928, 1006, 3073; Bradfield, Jones and Spencer, *ibid.*, 1931, 2907.

² Noelting and Collin, Ber., 1884, 17, 261

The following table shows the effect of salt formation upon the process of nitration.

TABLE IV 1

Nitration of 4-Chloro-dimethylaniline Cl- (3 1 N(CH3))

Per cent conc. H ₂ SO ₄ by Weight.	Proportion of 2- to 3- nitro compound obtained.
50	95: 5 (very slow reaction)
70	85:15
74.5	100: 0
83.9	100: 0
86∙o	92: 8
88.3-90	indefinite (oils obtained)
92.0	9:91
93·8 98·0	6 : 94
98.0	6:94

(d) The Directive Effect of Groups not Conjugated with the Aromatic Ring.

There is a very marked difference between the rate of substitution of the hetero-enoid substances that have been discussed in the preceding section and of all other aromatic compounds.

Of the less reactive substitution derivatives of benzene some, such as toluene, yield, on further substitution, mixtures of all the possible isomers in proportions which indicate that the energy required for the activation of the various carbon atoms does not vary within very wide limits, but others, such as nitrobenzene, yield one particular isomer in a disproportionally great amount (90% or more).

One may ascribe small differences in reactivity to differences in the intensity of local induced electrostatic fields (I or F), but the general effect of an induced electrostatic field of a substituent group does not produce any alternation in the degree of reactivity of successive atoms in any chain or ring (of. p. 239). Hence one must consider that the anionoid reactivity of particular carbon atoms in all benzene deriva-

¹ Clemo and Smith, J. Chem. Soc., 1928, 2414.

² Ingold and Shaw, f. Chem. Soc., 1927, 2918; Lapworth, Ingold, Rothstein and Ward, ibid., 1931, 1959.

tives appears only after electromeric change. All substituent groups reduce the symmetry of benzene, and consequently the probability of activation of each of the double bonds of a benzene derivative will be different.

Vorländer's generalization ² (p. 437) that positive groups are meta-directing and negative groups ortho—paradirecting is based upon an overwhelming mass of experimental fact, and hence the obvious approach to the question of what is the action of a substituent group is to consider the influence of an electrical field upon the conjugated ring of benzene.

Case I. The substituent is a negative pole, or is attached to the benzene ring at the negative end of a dipole, e.g.

$$\overrightarrow{CH_3} = \overrightarrow{C}$$

$$\overrightarrow{CH$$

The negative substituent will induce polarization in the aromatic ring in the direction indicated in Fig. XXVI above, and, in particular, the bond between carbon atoms r and z will be polarized so that, on activation, the negative pole will tend to develop at carbon atom z. (Compare Markownikow's rule, and the addition of HBr to CH_3 —CH— CH_4 , p. 178.)

¹ Robinson, Institute of Chemistry Lectures, 1932, p. 40, footnote; J. Society of Dyers and Colourists, 1934, 70.

² Ber., 1919, **52**, 263; 1925, **58**, 1893.

Now benzene is a conjugated system, and, by a process of electromeric change, the activated molecule (XXVIIA) could easily give rise to (XXVIIB). Activation of the double bond C_{\bullet} — C_{\bullet} is also possible (XXVIIc), but, as this bond is more distant from the polar substituent, its mode of activation will be less affected than all others, as it is situated in what must be a very weak electrostatic field.

All the carbon atoms in the aromatic ring of toluene are situated in an electronegative field, and consequently can all become anionoid centres more easily than can the carbon atoms in benzene. Hence, though the reactivity of the *ortho* and *para* carbon atoms will be particularly enhanced, the probability of substitution in *each* position is increased.

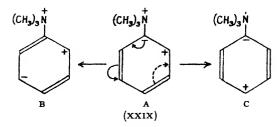
For substitution of toluene, using acetyl nitrate at 30° C., Ingold, Lapworth, Rothstein and Ward have found the probabilities of substitution of the *ortho*, *meta* and *para* positions to be 40.0:3.0:51 respectively, compared to benzene as 1.1

Case II. The substituent is a positive pole, e.g. $N(CH_3)_3$, or is attached to the benzene ring at the positive end of a dipole, e.g. $C \longrightarrow NO_2$.

The positive substituent will induce polarization in the aromatic ring in the direction indicated in Fig. XXVIII, and, in particular, the bond between carbon atoms will be polarized so that, on activation, the negative pole will tend to develop on carbon atom r, at which the substituent is attached.

450 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

By electromeric change in the benzene ring the activated structures (XXIXB) and (XXIXC) can be formed.



Hence a kationoid reagent would preferentially attack the meta position. Alternatively, the kationoid reagent might displace the substituent group by attacking carbon atom *t* of structure (XXIXA). Such group displacement is by no 1.1eans uncommon with meta-directing substituents, such as COOH, NO₂, or SO₂H, which Flürscheim regarded as "weakly attached" to the benzene ring (p. 429).

Fig. XXIX also indicates that anionoid reagents should tend to attack the kationoid carbon atoms of the *ortho* and *para* positions. This can, in fact, occur (see pp. 453-459).

Since amongst the *meta*-directing substituent groups are classified many groups which possess very large dipoles, e.g. the nitro-group, it is not surprising that one substitution reaction should prevail over all others. In the case of nitrobenzene the *meta*-directive power of the nitro-group may be enhanced still further by salt formation.²

The meta-directing substituents place every carbon atom of the aromatic ring in an electropositive field, and hence substitution by a kationoid reagent will be hindered generally. Hence, as Hollemann pointed out (p. 431), an ortho—para-directing group has always a more powerful orienting influence than a meta-directing group.

In conclusion one must emphasize the view that the induction of electromeric changes in an aromatic ring is due essentially to the *resultant* electrostatic field of the

¹ De Lange, Rec. trav. chim., 1926, 45, 19.

² Hetherington and Masson, J. Chem. Soc., 1933, 105.

substituent group. The predominating influence will be that of the single covalent bond by which the substituent group is attached to the nuclear carbon atom. The electrostatic field from this bond will be that from the vector sum of its own intrinsic dipole moment plus the dipole moment induced in it by more distant groups. The induced dipole moment is usually very small (cf. pp. 255-256).

A vast amount of recent experimental work, which cannot be discussed here, has been carried out with the object of comparing the relative directive powers of different substituent groups.¹ Their results serve to amplify the theory which has been described above, and no striking exceptions to the generalizations exist.

The dipole measurements of Sutton,² which have already received mention (pp. 255-256), indicate that the dipole moments induced in aromatic rings by substituent groups are, in fact, so oriented as to induce aromatic substitution of the types described in this section. Thus the induced dipole is positive for ortho — para-directive substituents and negative for meta-directive substituents.

TABLE V
Induced Dipole Moments of Aromatic Compounds

	-		
R—————	R——		
Ortho — para-directive groups. R Induced moment.	Meta-directive Groups, R Induced moment.		
CH ₃ + 0.45 OH + 0.23 NH ₂ + 0.32 Cl + 0.59 Br + 0.69 I + 0.88 CH ₂ Cl + 0.21	CCl ₃ - 0·5 COCH ₃ - 0·18 CHO - 0·28 CN - 0·43 NO ₂ - 0·90		

¹ For summaries, compare Waters, Chemical Reviews, 1930, 7, 407; Robinson, "Substitution in the Benzene Nucleus", J. Soc. Dyers and Colourists, 1934, 65.

² Proc. Roy. Soc., 1931, A 133, 668.

452 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

It is possible that the ortho — para-directive influence of the halogens may be due primarily to the inductive effect in the aromatic ring, and not to the conjugation with the nucleus as suggested on page 443. Halogen substituents are peculiar since they are ortho — para-directive and yet decrease the reactivity of the benzene ring. The positive end of the carbon—halogen dipole is directed towards the benzene ring, yet the dipole moment of chlorobenzene is less than that of methyl chloride.¹

¹ For discussion of this peculiarity, see: Ingold, Chemical Society Annual Reports, 1927, 24, p. 156; Shoppee, J. Chem. Soc., 1930, 974; 1932, 696; Baddeley and Bennett, ibid., 1933, 261; Bennett, ibid., 1933, 1112; Robinson, ibid., 1933, 1114; Ingold, ibid., 1933, 1120; Baddeley, Bennett, Glasstone and Jones, J. Chem. Soc., 1935, 1827; Sidgwick, J. Chem. Soc., 1936, 537.

CHAPTER XVII

AROMATIC COMPOUNDS—II

Kationoid Reactions of Aromatic Compounds.

In the previous chapter it has been shown that the normal process of substitution of aromatic compounds comprises the replacement of a hydrogen atom by a kationoid group, which attacks the aromatic ring at an anionoid carbon atom. This substitution is most facile with compounds which are hetero-enoid systems, for in these there can easily occur an electromeric change by which, in the activated molecule, a negative pole may be developed on a carbon atom in the *ortho* or *para* position to the substituent group from which the electromeric change is initiated. E.g.

The converse type of electromeric change is possible in conjugated systems which are *katio-enoid* (p. 392), and which therefore tend to attract electrons from the aromatic ring. For example, nitrobenzene can undergo electromeric change as follows:

Several other *meta*-directive substituent groups, e.g. the :CO group, are similar in character to the nitro-group: even those which are not conjugated directly to the aromatic

ring can *induce* a similar type of electromeric change within the benzenoid system (cf. p. 447).

One would therefore anticipate that aromatic compounds having meta-orienting substituents could be attacked in the ortho and para positions by anionoid reagents, and this is indeed the case. Nitrobenzene will react with strong alkali to give o-nitrophenol, and with piperidine in presence of sodamide to give p-nitrophenyl-piperidine, together with reduction products formed by the eliminated hydrogen. These instances of kationoid reactions can easily be explained by the electromeric oscillation sequence:

which as usual passes through an intermediate quinonoid phase. That quinonoid complexes such as (A) above can actually be formed by the action of alkalis on aromatic nitro-compounds has been proved by Meisenheimer,³ who obtained the same addition-product (B) from p-nitroanisole

¹ Wohl, Ber., 1899, 32, 3486.

² Bradley and Robinson, J. Chem. Soc., 1932, 1254.

³ Annalen, 1902, 323, 219, 241.

and sodium ethoxide and from p-nitrophenetole and sodium methoxide.

Polynitro-compounds are still more reactive; thus m-dinitrobenzene will react with potassium cyanide 1

$$\begin{array}{c}
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\
 & & \\$$

The manufacture of the dyestuff alizarin by fusing anthraquinone- β -sulphonic acid with caustic alkali and an oxidizing agent, such as a chlorate or a nitrate, affords another example of a reaction of an anion with a katio-enoid system (see top of page 456).

The carbonyl group promotes the electromeric change (a) (by which the carbon kation is produced) most easily when in alkaline solution (compare the conditions under which enolization occurs), and the elimination of the hydrogen is facilitated by the presence of the oxidizing agent which converts it to water. The conjugation of the carbonyl group with the benzene ring in anthraquinone is substantiated by the fact that the alizarin lakes are chelate ring compounds of considerable stability.

In the pyridine and quinoline series nuclear attack by an anionoid reagent is not at all infrequent. Thus sodamide easily converts pyridine to α -aminopyridine.²

The formation of a pseudo-base from a quinoline salt by

¹ Lobry de Bruyn, Rec. trav. chim., 1889, 2, 210.

² Tschitschibabin, J. Russ. Phys. Chem. Soc., 1914, 46, 1216.

the action of alkali ¹ involves, in the first instance, the neutralization, through an electromeric change, of the positive charge on the nitrogen atom of the pyridine ring, and is followed by combination of the kationoid carbon atom with a hydroxyl anion, the reaction being exactly similar in type to those discussed above (see top of page 457). With heterocyclic compounds the tendency to quinonoid (e.g. pyrone or pyridone) ring formation is so pronounced that the α - and γ -positions with respect to the hetero atom are rarely centres of anionoid polarization. Normal aromatic substitution is confined to the β -position, and, further, aromatic character is not fully displayed by substituents in α - or γ -positions. The hydroxy compounds, for example, are not phenolic, amino groups cannot be diazotized, and halogen compounds are easily hydrolysed.

¹ Decker, Ber., 1902, 35, 2588.

Pyrrole, in sharp contrast to pyridine, is very easily substituted by kationoid reagents in both the α - and the β -positions. Thus with iodine it yields tetra-iodopyrrole. Robinson has pointed out that the β -positions of pyrrole are analogous to the *ortho* positions of aniline or phenol, whilst the α -positions are analogous to a *para* position.

In benzene derivatives the kationoid substitution of hydrogen is not facile, for the atom has to be eliminated as an anion, H. This can only be accomplished successfully when there is also present an oxidizing agent that can accept its electrons.

However, the elimination from aromatic rings of sub-

¹ Institute of Chemistry Lectures, 1932, p. 43.

stituents such as the halogens, or nitro or sulphonate groups, is much more easy, since these radicals form stable anions, e.g. Cl⁻, NO₂⁻, just as hydrogen forms a stable kation. In accordance with this, one finds that the latter groups may readily be eliminated from katio-enoid substances. For example, 2:4-dinitrochlorobenzene contains a chlorine atom of "aliphatic character" which readily reacts with alkalis, ammonia or amines, e.g.

The *ortho* and *para*-halogenated benzoic acids also are reactive, though the reactivity is much less than that of the nitro-compounds and usually Ullmann's catalyst, copper powder, is required before the halogen replacement occurs at moderate temperatures.

Kationoid substitution of the type instanced above is a reaction of considerable commercial importance. Not only can the reactive halogen compounds be used for the manufacture of diphenylamine derivatives of importance to the dyestuff industry, but also the reaction can be adopted for the purification of aromatic nitro-compounds. In the mono-nitration of toluene, for instance, there is produced 4% of meta-nitrotoluene, which, in the later stages of trinitrotoluene manufacture, becomes converted to unsymmetrical nitrotoluenes, such as (I) on page 459. These not only lower the melting-point of 2:4:6-trinitrotoluene, but greatly decrease its stability to mild concussion, and also tend to decompose, in presence of moisture, with the production of corrosive acidic substances. Pure T.N.T., how-

ever, can easily be obtained by warming the crude nitration product with sodium sulphite, which with the unsymmetrical compounds yields sodium salts (e.g. II) easily soluble in water, but which leaves the symmetrical trinitrotoluene unattacked.

Another important kationoid substitution reaction is the decomposition of p-nitroso-dimethylaniline with hot alkali, whereby pure dimethylamine may be obtained.

Side-Chain Reactivity.

Electromeric changes of similar type to those dealt with above can lead to the activation of the side-chains of the aromatic nitro-compounds and of the heterocyclic bases. For instance, 2:4-dinitrotoluene has a reactive methyl group, capable of condensing with aldehydes and with nitroso compounds. It is, in fact, similar in character to the methylene group in keto-enolic substances.

The condensation between dinitrotoluene and benzaldehyde, which readily occurs in the presence of a trace of piperidine, or other strong organic base, can easily be explained by postulating a sequence of changes involving:

460 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

an electromeric change resulting in the liberation of proton from the methyl group, followed by:

the reversal of this electromerization, subsequent to the dissociation of proton, leading to the formation of an anionoid carbon atom at the α -position in the side-chain, which is then readily attacked by a kationoid reagent such as the carbon atom of an aldehyde.

$$\begin{array}{c|cccc}
\hline
CH_1 + CH - C_0H_5 & CH = CH - C_0H_5 \\
\hline
C & & C \\
HC & C - NO_2 & HC & C - NO_3 \\
HC & CH & HC & CH
\end{array}$$

The basic catalyst undoubtedly functions by promoting the electromerization of stage (1) above.

A similar reactivity of methyl groups is found in α - and in γ - methyl pyridine, and also in quinaldine and in lepidine.

The "Negative" Character of the Phenyl Group.

The sequence of electromeric changes described above can occur in any compound containing the benzyl group, C₀H₅—CH₂. Hence the CH₂ group in phenyl-acetic acid,

 C_0H_5 — CH_2 ·COOH, benzyl cyanide, C_0H_5 — CH_2 ·CN, desoxybenzoin, C_0H_5 — CH_2 —CO· C_0H_5 , and other substances, is replaceable by alkali metals. In consequence, the phenyl group is frequently regarded as a strong "negative", or acylous, group, and actually benzoic acid $(K = 7 \times 10^{-8})$ is a stronger acid than acetic acid $(K = 1.8 \times 10^{-8})$.

The acidity of benzyl cyanide and similar compounds must undoubtedly be due to the electromeric effect, which is, as shown above, reversible, it being possible for electrons to move both in to and away from the aromatic ring.

Thus, in addition to noticing the enhanced reactivity of the hydrogen atoms in benzyl cyanide, we find that the halogen atoms of the benzyl halides are easily liberated as halide anions. Dichlorodiphenylmethane, C_6H_6 — CCl_2 — C_6H_5 , hydrolyses with water alone, whilst triphenylchloromethane, $(C_6H_6)_3C$ —Cl, can exist in the form of a carbonium salt (p. 187).

The complete elucidation of the mechanisms of the reactions of aromatic compounds thus requires the recognition of the possibility of occurrence of reversible, or oscillatory, electromeric changes, in which electrons may either travel to, or be provided from, aromatic rings. Usually, conjugated substituent groups, which are either heteroenoid, such as —OR or —NR₂, or else katio-enoid, such as the nitro-group or the nitrogen atom in the pyridine ring, provide the origin or the ultimate seat of the electron movements. Occasionally, however, as in the examples cited above, the phenyl group itself provides or accepts electrons and so promotes the change. For this reason Ingold 2 has

¹ Cf. Ashworth and Burkhardt, J. Chem. Soc., 1928, 1791.

² Chemical Society Annual Reports, 1928, p. 147; J. Chem. Soc., 1929, 447; 1928, 904; Proc. Leeds Phil. Soc., 1, 421-31.

classified the phenyl group as a "reversibly tautomeric system" $(\pm T)$, and has ascribed the well-known experimental fact that aromatic rings are conducive to the stability of free hydrocarbon kations, anions, and free radicals to this property of being able to donate or absorb electrons, distributing the deficiency or excess of electrons throughout the aromatic ring system without appreciably weakening the conjugated ring structure. The stability of the aromatic ring, which has been ascribed to "resonance", thus extends, in part, to attached substituents.

Substitution of Condensed and Linked Aromatic Ring Systems.

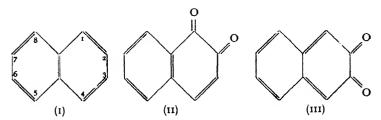
1. Naphthalene.

The substitution of any condensed ring system, such as that of naphthalene, is a very complex matter to elucidate, since, owing to the complete conjugation of the whole system, a much larger number of atoms is necessarily involved than in the case of benzene derivatives. Many important analogies, however, may be traced between naphthalene substitution and benzene substitution. In addition, further indications of the possible mechanisms of substitution processes can be gained from consideration of the known possibilities of benzenoid-quinonoid transformation within the molecule.

Naphthalene itself substitutes almost exclusively in the α -position at low temperatures; the only prominent exception to this rule being the β -sulphonation at 160°. In connection with this generalization may be recalled the fact that Thiele, in his original development of his theory of conjugation, pointed out that the deactivating effect of alternation of single and double linkages is not quite as complete in the α - as in the β -positions in naphthalene. It is also significant that α -substitution of naphthalene is often preceded by I: 4-addition. Thus when naphthalene is chlorinated there is produced, in the cold, naphthalene dichloride, $C_{10}H_{8}Cl_{20}$, which, on warming, loses hydrochloric

acid, yielding the more stable substitution product $C_{1\nu}H_{7}Cl.^{1}$

For the disubstitution of naphthalene one may state, as a general empirical rule, that if the first substituent is of an activating polar character, such as -OH, -NH, or CH₂, then the second substituent will enter the same ring as the first, but otherwise the reaction product will contain one substituent in each ring. Within a single ring the ordinary ortho - para-directive influence of an activating grouping seems to be exerted, but whilst an amino or hydroxyl group in the α -position will promote substitution in both the 4- and the 2-positions to it, a similar group in the β -(=2) position will only promote substitution in the adjacent α -(I) position, and not in the other β -position (3), to which it stands in ortho relationship equally. This significant fact has been cited as evidence that naphthalene has the symmetrical structure (I), and, in connection with it, one may point out that whilst I: 2-naphthaquinone (II) is known, its 2:3-isomer (III) is not.



When the disubstitution of naphthalene occurs in both rings it is found that, in general, substituents tend to occupy the α - rather than the β -positions, no matter what the polar nature of the first substituent. Thus, when the first substituent occupies position r, disubstitution in the positions 5- (the main product) and 8- is promoted equally by groups such as OH, NH₂, Cl, Br, CH₃, which are ortho—para-orienting, and by groups such as NO₂ and COOH which are meta-orienting in benzene. This may be taken

¹ Laurent. Annales de Chimie, 1835, 59, 196; 1837, 66, 196.

as indicative of the relative independence of action of the two rings in naphthalene, but, alternatively, may be considered as an effect of the polar action of the one conjugated system on the other, since this alone will lead to the activation of all the α -positions.

Polar activation in the naphthalene series can, however, be shown to be capable of transference from one ring to another, since, when an ortho — para-directive substituent occupies a β - (= 2) position in one ring, the diametrically-opposite 6-position may be the point of disubstitution. This is indeed well known in the sulphonation of both β -naphthol and β -naphthylamine, although these reactions are usually carried out at a fairly high temperature, under which condition β -positions always seem to be the more prone to sulphonation. Nitration of the β - chloro, bromo and alkoxynaphthalenes, however, takes the same course, the final product being the 2:6:8-trisubstitution derivative. This substitution is most probably one involving the electromeric oscillation sequence shown below,

which of course must occur in the production of amphinaphthaquinone.

Proof that such an electromeric change does occur in the naphthalene series during substitution reactions has been afforded by investigation of the bromination of β -naphthol. This reaction occurs as indicated on page 465.

¹ Van der Kam, Rec. trav. chim., 1926, 45, 564, 722.

² Fries and Schimmelschmidt, Annalen, 1930, 484, 245.

The quinonoid compounds contain reactive bromine atoms, capable of reduction with hydrogen iodide, but the benzenoid derivatives are stable in absence of strong acid.

2. Ouinoline.

Quinoline, as regards its substitution reactions, exhibits a close resemblance to α -nitronaphthalene, just as pyridine in many ways resembles nitrobenzene. Thus it nitrates in the 5- and 8- positions and not in the ring containing the strongly deactivating nitrogen atom. Electromeric changes which are so facile in the pyridyl ring do not seem to be transmitted to the benzenoid ring, although the general polar influence of the hetero atom deactivates the whole molecule.

4. Diphenyl.

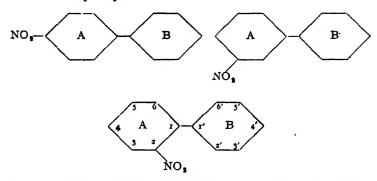
The linked ring system of diphenyl, which contains two independently complete benzene rings, seems to be one in

466 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

which electromeric changes do not tend to pass from ring to ring with any facility, although the whole molecule does possess an unbroken conjugated system throughout, and is capable of yielding binuclear quinonoid compounds by oxidation both of benzidine and of 4:4'-diphenol: e.g.



On substitution, however, not only diphenyl itself but also all substituted diphenyl compounds, $C_{\bullet}H_{\bullet}-C_{\bullet}H_{\bullet}X$, yield derivatives in which phenyl is substituted in the ortho and para positions to the linking bond. For example, all three nitrodiphenyls



give, on further nitration, only products in which the second ring (B) is substituted in the 4'- or the 2'- position.¹ The same rule applies for substitution of the aminodiphenyls and their derivatives. It is therefore evident that molecular activation by the electromeric process, which affects alternate positions in the molecule, cannot operate from one ring to another in the diphenyl system. Within a single ring, however, the normal laws of benzene substitution operate, and substitution can easily be effected in the ortho or para position to any activating substituent groups which may be present. Thus 4-amino-, 4-hydroxy-diphenyl and their derivatives readily undergo substitution in the

¹ Turner, Rec. trav. chim., 1929, 48, 821.

adjacent 3- position, and 3-aminodiphenyl and 3-methyldiphenyl are substituted in the 4- position.

In contrast, deactivating groups, such as the nitro-group or the halogens, inhibit aromatic substitution in the aromatic nucleus containing them, and consequently the second substituent enters the other ring. For example, 4-chloro-diphenyl gives, on nitration, a mixture of 4-chloro-2'-nitrodiphenyl and 4-chloro-4'-nitrodiphenyl

$$NO_{2}$$
 . $CI \longrightarrow CI \longrightarrow + CI \longrightarrow NO_{2}$

One has thus, with diphenyl, a compound which provides a means of ascertaining whether a substituent group promotes or hinders the occurrence of aromatic substitution.

Although induced electromeric activation from ring to ring is not possible in the diphenyl series, yet general polar (i.e. field) effects, which can be exerted alike in normal and in activated molecules, can be shown to operate throughout the whole of the diphenyl system.

In considering the substitution of a compound, $C_{\bullet}H_{\bullet}$ — $C_{\bullet}H_{\bullet}$:X, it can be considered that, although not only the unsubstituted phenyl group, $C_{\bullet}H_{\bullet}$:, but also all substituted phenyl groups, $C_{\bullet}H_{\bullet}$:X—, behave as strong ortho — para-directive groups, yet $C_{\bullet}H_{\bullet}$ X is a more powerful directing group than $C_{\bullet}H_{\bullet}$ — if X· is an activating, ortho — para-directive substituent such as NH_2 , OH or CH_3 , and a less powerful directive group than $C_{\bullet}H_{\bullet}$: if X· is a deactivating substituent such as NO_2 .

The action of a polar field in producing an alteration in the strength of the directive power 3 of a group without altering its directive type (e.g. from the ortho — paradirective class to the meta-directive class) can be paralleled in the toluene series, since one can greatly alter the nature

¹ Scarborough and Waters, J. Chem. Soc., 1920, 557; 1927, 89, 1133; Blakey and Scarborough, ibid., 1927, 3000.

² Grieve and Hey, J. Chem. Soc., 1932, 2245.

⁸ The directive power of a substituent group may be gauged by noting the place of the substituent in Hollemann's series (p. 431)

of a group ·X in the compound C_0H_0 — CH_2 ·X without altering the predominating ortho — para-directive action of the group — CH_2 . In the toluene series the action of the side-chain substituent X is revealed by measurement of the ratio

percentage yield of meta isomer percentage yield of (ortho + para) isomers

which does vary greatly (cf. p. 260).

Similarly, in the diphenyl series, a change in the polar field of the substituent group, $C_{\bullet}H_{\bullet}$:X, due to change in X, will alter the ratio between the yields of the possible isomers. Frequently this effect is sufficiently pronounced for the whole type of substitution reaction to be changed. Thus 4-hydroxydiphenyl is substituted first in the adjacent position 3, but its methyl ether gives, in addition, some 4'- substituted product, and, in the case of the acyl derivatives, the 4'- substituted product forms the main bulk of the yield.

$$HO \longrightarrow HO \longrightarrow NO_{1}$$

$$CH_{3}CO - O \longrightarrow CH_{3}CO - O \longrightarrow NO_{2}^{1}$$

In the case of 4-acetylaminodiphenyl also, bromination can occur in position 4', although chlorination seems to be confined to position 3.2

Particularly interesting results have been obtained by E. E. Turner and his colleagues ³ in studies of the nitration of 4:4'-disubstituted derivatives of diphenyl, with which unsymmetrical substitution can result. For example, 4:4'-dichlorodiphenyl gives initially 4:4'-dichloro-2-nitro-

¹ Bell and Kenyon, J. Chem. Soc., 1927, 3044.

² Scarborough and Waters, J. Chem. Soc., 1926, 557; Kenyon and Robinson, ibid., 1926, 1242, 3050.

³ Rec. trav. chim., 1929, **48**, 821 (this is a review of several earlier papers).

diphenyl, and, on further nitration, this yields 4: 4'-dichloro-2: 3'-dinitrodiphenyl. E.g.

$$\begin{array}{c|c} Cl & Cl & Cl & NO_2 \\ \hline \\ \hline \\ Cl & Cl & Cl & Cl \\ \end{array}$$

This sequence indicates that one must place the relevant directive groups in the order

$$Cl \cdot C_6H_4 -> Cl -> Cl \cdot C_6H_3 \cdot NO_2 -$$

which indicates that the nitro-group generally hinders aromatic substitution by a kationoid reagent.

5. Other Linked Ring Systems.

The orientation effects that have been described for the diphenyl series are found to a greater or less extent in other linked aromatic systems of the general type C₅H₅—XY—C₅H₅, and vary according to extent to which the linking group -XY- permits the transmission of general induced polarization influences.1 The general conclusion to be drawn from a review of all such systems is that an aromatic ring preserves not only its own typical modes of substitution, as in benzene itself, but is, however substituted, always an ortho - para-directing substituent group, although its directive power may vary. It would seem that ortho - para-orienting properties are to be associated more particularly with the electronic type of structure possessed by benzene than with the particular atoms which constitute it, since even the three isomeric phenyl-pyridines substitute predominantly in the ortho and para positions to the pyridyl ring,2 although the nitrogen

¹ Waters, Chemical Reviews, 1930, 7, 406.

² Forsyth and Pyman, J. Chem. Soc., 1926, 2912.

atom in the pyridine nucleus has a most decided metaorienting property.

The extent to which general polar effects of substituents can operate in linked ring systems depends very largely upon the nature of the linking group. With strongly polar groups, such as >CO, there is but little evidence to suggest that a substituent in one ring can affect substitution in the other, but with less polar groups such as CH₂ or —O—these effects are more marked. Substitution of linked ring systems is, however, so complex that only the guiding principles can as yet be cited, and, in fact, the experimental work has not yet passed the early qualitative stage.

Some Reactions of Quinones.

p-Benzoquinone is a discordant katio-enoid system (p. 393) since the two carbonyl groups have opposing polarizing influences upon the ·CH=CH· group which separates them.

This effect of "crossed polarity" has already received some discussion with reference to maleic anhydride (p. 393), which can be regarded as the quinone derived from furane.

The substitution reactions of the quinones are unlike those of benzenoid rings, but show plainly that quinones must be regarded as $\alpha\beta$ -unsaturated ketones,—CO-CH=CH—. This similarity becomes most obvious when the symmetry of the quinone ring is destroyed by the introduction of a substituent.

With the halogen acids, acyl chlorides and acid an-

hydrides, quinones react to give derivatives of quinol, $C_0H_4(OH)_2$, and the reaction has frequently been thought to involve a reduction. It is, however, much more easily explicable as an addition-reaction. Thus Thiele 2 showed that p-benzoquinone would react with hot acetic anhydride containing sulphuric acid to give I: 2: 4-triacetoxybenzene, and the reaction may be represented quite simply as follows:

Erdtman ³ has shown that the ease of effecting this reaction is greatly influenced by the presence of substituents. For example, methoxy-p-benzoquinone (I) forms a neutralized conjugated system (a) with one of the carbonyl groups, rendering the remainder of the molecule a simple katio-enoid system (b) scarcely affected by crossed polarity. Hence the Thiele acetylation readily proceeds, with formation of (II). 2:5-Dimethoxyquinone (III), in contrast, is a completely neutralized system, inert towards acetic anhydride

¹ Cf. Meyer-Jacobsen, "Lahrbuch", 2, 1, p. 442.

⁸ Annalen, 1900, 311, 349.

⁸ Proc. Roy. Soc., 1933, A 143, 177.

Toluquinone (IV) is very similar to methoxyquinone, but far less reactive; it is, however, more reactive than p-benzoquinone. This is to be expected, as the methyl group can only act by *inducing* the occurrence of the elec-

tromerization (a), but does not form part of the conjugated system. p-Xyloquinone (2:5-dimethylquinone) will react slowly with acetic anhydride.

Quinones readily polymerize, and, on oxidation in alkaline solution, give brown amorphous products of similar nature to the humic acids obtained by the decomposition of plants. It has been shown that the first stage in this polymerization consists in the union of two quinone ring systems with formation of a derivative of the diphenyl series.

A similar carbon to carbon union occurs in the Diels reaction (p. 376), which can be carried out as easily with p-benzoquinone as with maleic anhydride. Robinson has interpreted this reaction as a further instance of addition to an unsaturated ketone. E.g.---

¹ Erdtman, Proc. Roy. Soc., 1933, A 143, 177-241.

² 4th Solvay Report, 1931.

(a)
$$HC$$
 CH_z
 HC
 CH_z
 HC
 CH_z
 CH_z

The electromeric change of stage (a), which leaves a kationoid carbon atom at position 3, is reversed in stage (b), which leaves an anionoid carbon atom at position a. The butadiene is alternately polarized, and reacts in stage (a) at its anionoid carbon atom, and in stage (b) at its kationoid carbon atom. The reaction with p-benzoquinone occurs twice, with the eventual production (c) of octahydroanthraquinone.

Many of the reactions by which azine dyestuffs may be made, by the condensation together of aromatic bases upon oxidation, are at bottom katio-enoid reactions of quinones.

Molecular Rearrangements of Aromatic Compounds.

1. Benzenoid-Quinonoid Changes.

Isomeric change between benzenoid and quinonoid structures has been recorded only with those molecules which are extended conjugated systems; e.g.—

474 PHYSICAL ASPECTS OF ORGANIC CHEMISTRY

(a) Nitrosophenol \rightleftharpoons Quinone oxime (page 298).

(b)
$$\begin{array}{c} & & & & \\ & & & & \\ & & & \\ N & & & \\ & & & \\ \alpha\text{-Hydroxypyridine} & & \rightleftharpoons & \alpha\text{-Pyridone} \end{array}$$

These changes, and many others of great technical importance, such as those of the phthaleins and of the triphenylmethane dyes, are typical ionotropic changes, in which the mobility of the migrating ion is facilitated by the fact that the two alternative structures have comparable stabilities. They are invariably reactions initiated externally by the use of reagents which can effect dissociation of the mobile group. Thus, the mobility of the phthaleins, citable as an example of prototropy, is brought about by the use of alkali, which promotes the ionization of hydrogen, whilst the isomerizations of the triphenylmethane dyes, which afford excellent examples of anionotropic changes, are brought about by the use of a mineral acid. Frequently these reactions proceed sufficiently slowly for one to be able to measure their reaction velocities by observations of colour changes; in these cases it can be demonstrated that the acidic or alkaline solvent acts as a catalyst in the change. The above reactions, therefore, differ in no way from those discussed in Chapter XIII.

2. Migrations of a Group from the Side-chain to the Nucleus of an aromatic compound occur only with those substances which have already been classified (pp. 442-446) as hetero-enoid systems.

The actual migration of a radical from a side-chain substituent group to a position directly attached to the nucleus, which is necessarily accompanied by the reverse migration of a hydrogen atom, is undoubtedly an ionic reaction, since it can be catalysed by an ionizing reagent.

Thus, the isomerizations of substances of the following types are all effected by the use of a strong acid, which, incidentally, converts a feebly basic starting product into a more basic end-product.¹

- (a) Tertiary aromatic amines.2
- (b) Nitrosamines.3
- (c) Nitramines.4
- (d) Chloramines.5
- (e) Sulphamic acids.6
- (f) Diazoamino-compounds.
- (g) Phenol ethers and esters.7

In each of these reactions the nuclear substitution by the migrating radical undoubtedly involves substitution by a kationoid group at an anionoid carbon atom, since the migrant radical enters the aromatic ring exclusively in the ortho or para position to the fixed substituent atom (N or O) which, with the aromatic ring, forms a conjugated hetero-enoid system. As in the case of a simple ionotropic change, the migrating group shifts from one atom to an alternate one along the conjugated chain.

The ionic type of the isomeric change cannot therefore be doubted, but the actual isomeric change may be complex in mechanism, and may involve two separate, though consecutive, reactions, since it has been pointed out 8 that the reagents used to effect the above isomerizations produce, or at least can produce, stable products capable of reacting directly with the aromatic nucleus. Further, the intermediate product is formed under exactly those experimental conditions in which direct substitution of benzene derivatives can be carried out. Thus, chlorine is produced by the action of hydrochloric acid on a N-chloroamine, and

¹ For a detailed review of these reactions, see Porter, "Molecular Rearrangements"; also Henrich, "Theories of Organic Chemistry".

² Hofmann and Martius, Ber., 1872, 4, 742.

³ O. Fischer and Hepp, ibid., 1886, 19, 2991.

⁴ Bamberger, ibid., 1894, 27, 584.

⁵ Chattaway and Orton, J. Chem. Soc., 1901, 79, 277.

⁶ Bamberger, Ber., 1897, 30, 654.

⁷ Claisen, Annalen, 1919, 418, 69; Claisen and Tietze, Ber., 1925, 58, 275; 1926, 59, 2344.

⁸ Cf. Michael, Ber., 1881, 14, 2107; J.A.C.S., 1920, 42, 787.

an acyl chloride by the action of anhydrous zinc chloride on diacetanilide. Again, alkyl groups may be made to migrate from side-chain to nucleus under exactly those conditions under which the Friedel-Crafts synthesis can be carried out.

Isomeric changes in which the intermediate products have a free life sufficiently prolonged for their molecules to be able to separate by kinetic migration in solution are said to be *intermolecular reactions*. In contrast, those reactions in which the intermediate entities have no perceptible free existence are termed *intramolecular reactions*. The reactions of Chapter XIV are all of this latter type.

Examples of Intermolecular Rearrangements.

Several isomerization reactions of aromatic compounds have been shown to be *inter*-molecular in type.

Thus, free diazobenzene chloride must be produced in the isomerization of diazoaminobenzene by the action of hydrochloric acid. If diazoaminobenzene is heated with hydrochloric acid and dimethylaniline there is produced dimethylaminoazobenzene, which can result only if the primary decomposition (A) is a reaction yielding products having independent existence, i.e.

A
$$\longrightarrow$$
 NH-N=N- \longrightarrow + HCl \longrightarrow -NH₁ + Cl-N=N- \longrightarrow (CH₂)₂N- \longrightarrow -N=N- \longrightarrow irreversible

H₂N- \longrightarrow + Cl-N=N- \longrightarrow H₂N- \longrightarrow -N=N- \longrightarrow

Similarly, N-chloro-acetanilide is converted into orthoand para-chloro-acetanilides only in the presence of hydrochloric acid.¹ Other transforming reagents, such as dilute acetic acid, act only by hydrolysing the N-chloramine to give some hydrochloric acid.² The hydrochloric acid acts

¹ Armstrong, J. Chem. Soc., 1900, 77, 1047.

² Orton and Jones, J. Chem. Soc., 1909, 95, 1456.

by producing free chlorine, which can interact either with the side-chain or with the aromatic ring.¹

$$C_{e}H_{s}\cdot NH\cdot COCH_{3}+Cl_{2} \underbrace{ \begin{array}{c} \mathbb{Z} & C_{e}H_{s}\cdot N(Cl)\cdot COCH_{3}+HCl \\ \\ & Cl\cdot C_{e}H_{4}\cdot NH\cdot COCH_{3}+HCl \\ \end{array} }_{}$$

The free chlorine can be removed by bubbling air through the transforming mixture, and the rate of formation of the nuclear chloro compound is equal to that in a corresponding mixture of acetanilide and free chlorine.² Further, the transformation yields a mixture of *ortho*- and *para*-substitution derivatives, identical in composition with that prepared by the direct chlorination of acetanilide.³

N-chloro-amines can, in fact, be used very conveniently as mild generators of chlorine, and several substitution reactions that are impracticable when free chlorine gas is used as the reagent can be carried out with their use. A free phenol or aromatic amine, for example, can be chlorinated smoothly by the use of a N-chloro-amine in dilute acetic acid.⁴

By the use of this reaction the actual rates of the constituent reactions in the chloro-amine transformation can be measured, and the influence of the catalysing acid can be followed quantitatively.⁵ The simultaneous reactions of N-chlorination (which is reversible) and C-chlorination (not reversible) are differently affected by the acidity of the medium. In glacial acetic acid the N-halogen compound cannot exist, but immediately decomposes to give free chlorine. As the acetic acid is progressively diluted with water the N-halogen compound becomes more and more stable, until with 50% acetic acid it is very rapidly formed from acetanilide and free chlorine. The speed of decomposition of N-chloro-acetanilide is, however, increased by increasing the acidity of the medium, and, in 1922, it was

¹ Orton, British Association Reports, 1910, 85.

² Soper, J. Phys. Chem., 1927, 31, 1192.

³ Orton and Bradfield, J. Chem. Soc., 1927, 986; 1928, 351

⁴ Orton and King, J. Chem. Soc., 1911, 99, 1185.

⁵ Orton, Soper and Williams, J. Chem. Soc., 1928, 998.

suggested by Harned and Seltz 1 to be proportional to the hydrogen ion activity of the solution. This would imply that the reaction by which N-chloro-acetanilide is decomposed is:

$$C_{\bullet}H_{\bullet}\cdot N(Cl)\cdot COCH_{3} + \overset{+}{H} \longrightarrow C_{\bullet}H_{\bullet}\cdot NH\cdot COCH_{3} + \overset{+}{Cl}$$

The reaction, however, requires also the presence of chloride anions, for free chlorine molecules, and not molecules of hypochlorous acid, are produced as the N-chloro-amine decomposes. In reality, the catalyst is unionized hydrogen chloride,² and, in an aqueous solution, the rate of reaction is directly proportional to the concentration of the hydrochloric acid molecules which are present.

The final ring-substituted product of an intermolecular isomerization of an aromatic compound may not have any direct structural connection with the starting product. Thus whilst phenyl-iododichloride (I) gradually transforms to p-chloro-iodobenzene, with liberation of free hydrogen chloride, the dichloride of p-iodophenol, (II), rearranges to give an *ortho*-chlorophenol, p the chlorine migrating to the *meta* position to the iodine atom—a position which it could not reach by any simple ionotropic change. The product (III) is actually that of the reaction between p-iodophenol and free chlorine.

The molecular rearrangements of the sulphonic acids of the naphthalene series seem to be rather exceptional also, in that the substituent group can migrate from one position to an adjacent one. If naphthalene is sulphonated there is produced a mixture of α - and β -sulphonic acids. Below

¹ J.A.C.S., 1922, 44, 1475.

² Dawson and Millet, *J. Chem. Soc.*, 1932, 1920.

⁸ McCombie and Brazier, J. Chem. Soc., 1912, 101, 968.

80° C. the α -acid forms the main yield, but above 100° appreciable quantities of the β -acid form, and at 160° or over this forms the main yield.¹ Further, on heating α -naphthalene-sulphonic acid in concentrated sulphuric acid, one obtains the same mixture of isomers as that which is produced by the direct substitution reaction.

Hence

Similarly, naphthalene 2:7-disulphonic acid gives, on heating with sulphuric acid, the 2:6-disulphonic acid.²

$$HO_3S$$
 \rightarrow
 HO_3S
 \rightarrow
 HO_3S

This is an isomerization reaction which may involve a reversible substitution, since sulphonic acid groups are easily eliminated from the aromatic nucleus. By the action of superheated steam, or by the action of water under pressure, it is possible to regenerate the hydrocarbon, and, in the case of the naphthalene sulphonic acids, the α - acid is much more easily decomposed in this way than the β-acid. Noelting 3 therefore suggested that the transformation could simply be explained by supposing that, at temperatures of 150° or above, a sulphonic acid group tended to hydrolyse from an α- position, and that the sulphuric acid produced thereby would react at a β - position with the naphthalene molecule. In elaboration of this, Erdmann 4 pointed out that if any disulphonic acid were temporarily formed in a reaction mixture then the percentage of water in the system would thereby increase, and, in

¹ Merz, Zeit. für Chemie, 1868, 393; Merz and Weith, Ber., 1870, 3, 195.

² Ebert and Merz, Ber., 1876, 9, 592; Heid, J.A.C.S., 1927, 49, 844.

⁸ Ber., 1875, 8, 1095. ⁴ Annalen, 1893, **275**, 192.

consequence, the hydrolysis of the α -sulphonic acid would be favoured.

The migration of sulphonate groups from α - to β - positions at high temperatures has been noticed also in the case of the acids derived from both the α - and the β - naphthylamines. In the case of β -naphthylamine, Green and Vakil have shown that the 2:8- and 2:5-sulphonic acids are formed below 100°, whilst the 2:6- and 2:7- sulphonic acids, containing only a β -sulphonate group, are formed at 150° or above. At the high temperatures the two former acids undergo a slow isomerization in presence of sulphuric acid. Further, at 160° or above, the 2:8- and 2:5-sulphonic acids hydrolyse to an appreciable extent in sulphuric acid of from 75% to 95% purity, whilst the 2:6- and 2:7-acids do not hydrolyse under these conditions. I.e.

The isomerization of the sulphonic acids of the naphthalene series, therefore, seems to be a case of reversible substitution, possible owing to the special ease of elimination of the $-SO_3H$ group from the aromatic ring.² One can regard the elimination of the sulphonate group from an α -position as an instance of replacement of an anionoid radical by hydrogen kation.

It may be pointed out that the α -positions in naphthalene are primarily centres of anionoid reactivity, and hence the activation indicated below would render naphthalene- α -sulphonic acid unstable:

¹ J. Chem. Soc., 1918, 113, 35.

² Cf. Euwes, Rec. irav. chim., 1909, 28, 298.

In connection with this interpretation of the migration of sulphonic groups one may instance the ease with which the SO₃H group may be displaced from aromatic rings by other groups (cf. p. 450). Thus, sulphanilic acid, on treatment with bromine water, yields tribromoaniline, whilst phenol-sulphonic acid on nitration gives picric acid.

Examples of Intramolecular Rearrangements.

By no means, however, are all cases of the migration of a radical from the side-chain to the nucleus of an aromatic compound capable of satisfactory explanation on the intermolecular reaction hypothesis.

For instance, the molecular rearrangement of the nitroamines differs in many respects from the rearrangement of the chloro-amines (p. 477). It is not affected by specific catalysts, for all acids are equally effective, and, moreover, no specific nitrating agent can be discovered in the isomerizing system.¹

Further, the Hofmann transformation of methylaniline hydrochloride to toluidine hydrochlorides can be carried out successfully by heating in an open tube at 300°, in spite of the fact that methyl chloride is a gas at room temperature.²

Similarly, the benzidine transformation can be brought about by the use of any mineral acid, and does not give rise to mixed products.³

Hence it is quite evident that an ion liberated from the side-chain of an aromatic compound can react immediately with the nucleus, and not combine with an ion of the transforming reagent. This is quite possible, since it is evident,

- ¹ Bradfield and Orton, J. Chem. Soc., 1929, 915.
- ² Hickinbottom, *ibid.*, 1934, 1700.
- ³ Ingold and Kidd, ibid., 1933, 984.

from the electromeric theory of aromatic substitution that has been discussed in the preceding pages, that an electromeric change leading to activation of the benzene ring can immediately follow any ionic activation of the side-chain which yields a free electron pair. Consequently a kation liberated from a side-chain substituent could react immediately with the nucleus without having any independent existence of appreciable duration. E.g.

There is definite evidence that the molecular rearrangement, with mineral acid, of phenyl-hydroxylamine to p-amino-phenol is a reaction of this type, since Bamberger found that para-substituted phenyl hydroxylamines formed initially quinoles, which then could re-form benzene derivatives by migration of either the hydroxyl or the alkyl group (R).1

By the action of heat, or with dilute aqueous sulphuric acid or dilute alkali, the alkyl group migrates, but with alcoholic sulphuric acid the hydroxyl group migrates.

¹ Bamberger, *Ber.*, 1894, **27**, 1349; 1900, **33**, 3600; 1902, **35**, 3886; 1903, **36**, 1625, 2028.

Similar quinonoid compounds can be obtained during the course of the Reimer-Tiemann reaction between p-cresol and chloroform. Quinonoid products can also be obtained in certain bromination reactions of phenols.¹

In conclusion, one can safely ascribe all migrations of groups from side-chains to nuclei of aromatic compounds to an ionic reaction in which the migrant group separates as a kation, leaving an unshared electron pair on the aromatic residue. The aromatic residue thereby becomes capable of, and may be stimulated to, an electromeric change, which can activate carbon atoms in *ortho* or *para* positions to the substituent group. These activated carbon atoms are anionoid in character.

The migrant kation, however, can either react with the active anionoid carbon atom of the aromatic nucleus or with any other anion in the reagent mixture which can donate to it an electron pair. Both spatial factors, energy contents, and the mean free lives of the ions need be taken into account before the exact sequence of the molecular rearrangement can be determined. A migration to an ortho position can obviously occur more easily than a migration to a para position, for in the latter change a large distance has to be traversed. One can, therefore, suggest that those kations which migrate to para positions must necessarily have a fairly long free life.

As yet one cannot estimate either the probability of occurrence, or the approximate duration, of an electromeric change in an aromatic ring. It is quite possible that the nuclear attachment of the migrating kation may require a different electromeric activation from that which could result from its initial expulsion from the side-chain, and that this second electromeric change may be brought about by a collision of the reactive kation with an appropriate atom in the aromatic nucleus.

¹ Auwers, Ber., 1902, **35**, 443, 465, 4207; 1903, **36**, 1801, 3002; Zincke, ibid., 1895, **28**, 3121; Annalen, 1901, **320**, 145; 1902, **322**, 174, 1904, **330**, 61; **334**, 342, 367.

AUTHOR INDEX

Adams, F. H., 340 Adams, R., 342, 343, 345, 346 Adkins, H., 135 Albrecht, H., 339, 343 Alder, K., 376, 414 Allan, J., 390, 392, 438, 444 Allsopp, C. B., 130, 397, 398, 400 Ampère, 115 Anderson, J. A., 59 Angeli, A., 110 Armstrong, E. F., 308 Armstrong, H. E., 426, 427, 433, 476 Arrhenius, S., 5, 81, 82, 93, 273, 274 Ashworth, F., 461 Auwers, K. von, 483 Avogadro, A., 54, 117 Bäckström, H. L. J., 90, 151 Baddeley, G., 237, 249, 452 Baeyer, A. von, 102, 188, 297, 298, 317, 408, 418 Bailey, C. R., 422 Baker, J. W., 249, 260, 299, 310, 318, 346, 350, 393 Ballauf, F., 122 Bamberger, E., 147, 299, 409, 410, 475, 482 Banfield, F. H., 125 Banta, C., 229 Barker, E. F., 27, 66 Barnett, E. de B., 435 Bates, J. R., 143 Beckmann, E., 80, 126, 325, 351-4 Behrend, R., 313 Bell, F., 342, 468 Bennett, G. M., 157, 235-7, 249, 452

Béraneck, J., 143 Berger, G., 230, 235, 236, 263 Bergmann, E., 127, 157 Bernal, J. D., 37, 406 Berry, W. A., 235-7 Berthollet, C. L., 296 Berthoud, A., 139–43, 150 Berzelius, J. J., x, 1, 2, 8, 11, 115, 116, 159, 288, 425 Bettag, L., 230, 231 Biilmann, E., 268 Birge, R. T., 14 Bischoff, C. A., 341 Bishop, G., 353 Bjerrum, N., 199, 214, 217, 280, Blakey, W., 284, 286, 287, 467 Blanksma, J. J., 428, 429 Bock, L. H., 342 Bodenstein, M., 49, 85, 132 Boeseken, J., 231 Bohr, N., 13, 126 Boltzmann, 250 Bone, W. A., 284, 426 Bonhoeffer, K. F., 119 Book, G., 94 Boorman, E. J., 179 Born, M., 199 Boullay, 116 Bowen, E. J., 45, 46, 150 Boye, E., 124, 157 Boys, S. F., 329 Bozorth, R. M., 35 Bradfield, A. E., 246, 445, 446, 477, 481 Bradley, R. S., 83 Bradley, W., 182, 369, 454 Brady, O. L., 353, 361, 362 Bragg, W. L., 19 Brazier, S. A., 478 Brigl, P., 309

Britton, H. T. S., 196 Broadway, L. F., 120 Brockway, L. O., 113, 407 Brönsted, J. N., 197, 198, 200, 313, 316 Brown, A. Crum, 137, 426, 427 Bruce, J., 373 Brühl, J. W., 401, 402 Brunner, M., 151 Bruyn, Lobry de, 265, 455 Bunsen, R. W., 117 Burgess, H., 37, 306 Burkhardt, G. N., 461 Burton, H., 157, 176, 177, 220, 224, 319, 367, 375, 396 Butlerow, A., 290-8, 317

Cady, W. G., 59 Cannizzaro, S., 117 Carlisle, 1 Carothers, W. H., 163 Cashmore, A. E., 265 Caspari, W. A., 216 Cathala, J., 88 Chapman, A. W., 157, 321, 322, 356-9 Chattaway, F. D., 475 Christiansen, J. A., 90, 91 Christie, G. H., 341 Claisen, L., 301, 475 Clar, E., 131, 414 Clark, W. M., 196 Claus, A., 7, 420 Clausius, R., 54, 59 Clemo, G. R., 447 Clough, G. W., 327, 328 Coffmann, D. D., 163 Collie, J. N., 418 Collin, 446 Colonius, H., 192 Conant, J. B., 124, 226 Cone, L. H., 191 Connor, R. J., 228 Cook, J. W., 435 Crafts, J. M., 230 Crum Brown, A., 137, 426, 427 Curtius, T., 111, 147, 363, 364 366, 367, 368

Dadieu, A., 317 Dale, 400

Dalton, J., x Darapsky, A., 146 Darzens, G., 230, 231 Davy, Sir H., 115, 116, 195 Dawson, H. M., 270, 274-7, 478 Debye, P., xiii, 22, 51, 52, 54, 55, 58, 63, 71, 245, 250, 251 Decker, H., 389 De Lange, M. P., 429, 432, 450 De Laslo, H., 14, 110, 406 Delbrück, 309 Dewar, Sir J., 408, 420 Diels, O., 376, 377, 414, 472 Dougherty, G., 370 Drugman, J., 426 Dubrunfaut, 306 Dufraisse, C., 90, 150 Dumas, J. B., x, 2, 116, 159 Dunn, F. P., 361, 362 Duppa, 300

Ebert, G., 479 Ebert, L., 58, 59, 60, 78 Edlund, K. R., 26, 172, 173 Egerton, A. C. G., 90, 152 Eggert, J., 94 Egli, H., 241 Ehrhardt, F., 71 Eichelberger, L., 175 Einstein, A., 85, 143, 149 Eisenlohr, 400 Emschwiller, G., 143 Engel, H., 435 Erdmann, H., 312, 479 Erdtman, H. G. H., 471, 472 Erlenmeyer, E., 6, 378 Errera, J., 60, 64 Estermann, I., 51, 60, 61, 64 Eucken, A., 69 Euler, H., 159 Euwes, P. C. J., 480 Evans, J. T., 375 Evering, B. L., 145 Ewald, L., 124

Fajans, K., 20, 22, 51, 198, 202, 203, 251, 257 Faraday, M., 420, 429, 430 Farkas, L., 86 Farmer, E. H., 373, 375, 377 Faulkner, I. J., 314

Fischer, Emil, 307, 309, 327, 331, Fischer, O., 475 Fittig, R., 135 Fleming, G. H., 367 Flürscheim, B., 156, 157, 219, 238, 429, 430, 436, 450 Forster, M. O., 364 Forsyth, R., 469 Fothergill, R. E., 165 Fowler, R. H., 37 Francis, A. W., 174, 176, 429, 432 Franck, J., 86 Frankenburger, W., 148 Frankland, P., 300 Frankland, P. F., 332 Franzen, H., 244 Fraser, R., 120 French, H. S., 399 Freudenberg, K., 58, 327 Friedel, C., 230 Fries, K., 435, 464 Fry, H. S., 160, 161, 388, 418, 419, 427, 435, 436

Gadamer, J., 332, 333, 334, 338 Gane, R., 215-17, 281 Garner, W. E., 332-4 Gates, S. F., 90 Gay-Lussac, 115, 116, 288 Gautier, 318 Gerhardt, K., 3 Gerlach, W., 60, 120 Geuther, 300 Ghosh, J. C., 144 Gibby, C. W., 265 Gibson, J., 426, 427 Gilman, H., 165, 192 Gladstone, 400 Glasstone, S., 138, 452 Goldberg, A. A., 375 Goldschmidt, H., 272 Goldschmidt, S., 126 Gomberg, M., 117, 118, 122-4, 145, 156, 187, 191 Gottlieb-Billroth, H., 124 Goudsmit, S., 16 Green, A. G., 480 Gregory, D. V., 230 Grieve, W. S. M., 147, 467

Grignard, V., 164, 165, 186, 225 Griner, G., 372-3 Groger, R., 111, 112 Groocock, C. M., 279 Guggenheim, E. A., 313, 316

Haber, F., 86 Hambly, F. J., 289 Hamilton, E. E. P., 170 Hammick, D. L., 133, 427 Hantzsch, A., 30, 36, 100, 166, 187, 198, 205, 304, 352 Harned, H. S., 478 Harries, C., 374, 409 Harteck, P., 86, 119 Hartel, H. von, 133 Heid, J. L., 479 Heitler, W., 15, 46 Helfenstein, A., 382 Helmholtz, ix Henderson, S. T., 317 Henderson, T., 168 Hendricks, S. B., 112 Henri, V., 14, 87 Henrich, F., 142, 160, 206, 212, 475 Hepp, P., 475 Hermann, 304 Hertzberg, G., 46 Herzenstein, A., 123 Hetherington, J. A., 450 Heurlinger, 14 Hey, D. H., 147, 218, 467 Hickinbottom, W. J., 481 Hickling, A., 138 Hildebrand, J. H., 61 Hinshelwood, C. N., 83, 84, 91. 246, 247, 271, 285 Hirst, E. L., 168 Hissinger, 1 Hofeditz, W., 119 Hofer, H., 138 Hofmann, A. W., 363, 364, 366, 367, 368, 370, 408, 475 Hollemann, A. F., 285, 429-44, 450, 467 Holtz, J., 192 Homeyer, A. H., 365 Hoogewerf, S., 366 Hopff, H., 318 Howis, C. C., 356

400 11110101111	
Hubner, 425	Kenyon, J., 125, 126, 328, 329,
Hückel E 22 20 155 158 150	336, 338, 342, 468
Hückel, E., 22, 39, 157, 158, 250,	Kermack W () 20 220 218
251, 396, 422	Kermack, W. O., 39, 220, 348,
Hückel, W., xi, 231, 382, 410, 412	390, 391, 438
Hudson, C. S., 312	Kestner, 288
Huggins, M. L., 37	Ketelaar, H., 60
Hughes, E. D., 84, 246, 247, 341	Kharasch, M. S., 172, 179
Hund, F., 17	Kidd, H. V., 481
Hunter, R. F., 18, 35, 41, 42, 48	Kindler, K., 286
Hussey, R. E., 226	King, H., 477
•	Kipping, F. S., 101, 306
Illingworth, W. S., 427	Kirkbride, F. W., 144, 148
Ingold, C. K., 76, 84, 114, 130,	Klinger, 374
157, 158, 161, 176, 177, 180,	Knaggs, I. E., 113
192, 215-17, 220, 234, 240-	Knorr, C. A., 203, 257
4, 246, 247, 253, 260, 269,	Know I 203, 237
277, 279-82, 299, 304, 305,	Knorr, L., 301
310, 348, 349, 375, 388, 393,	Kobeko, P. P., 59
394, 396, 408, 421, 423, 432,	Kohn, O., 127
	Kolbe, H., 137
438, 440, 441, 447, 449, 452,	Kon, G. A. R., 326
461, 481	Kondatow, 231
Ingold, (Mrs. E. H.), 110, 161,	Kossel, W., x, 2, 10, 19
234, 388, 438, 440	Kramers, H. A., 91
Iredale, T., 143	Kraus, C. A., 118, 193
Irvine, J. C., 309	Kronig, L., 43
Jackson, A., 279	Kuhara, 354
Jackson, E. L., 176	Kuhn, R., xiv, 48, 120, 128-30,
Jackson, L. C., 120	327, 339, 343-5, 382, 397, 412
	Kurtshatow, I. V., 59
Jacobsen, P., 303, 304, 310, 405,	
471 Jonling H. O. 48	Laar, C., 297, 298, 300, 302-5, 316
Jenkins, H. O., 58	Ladenburg, A., 405
Job, A., 143	Landolt, 307, 400, 408
John, F., 131	Langevin, 54
Johnson, G., 135	Langmuir, I., 11, 12, 28, 43, 45,
Jones, B., 237, 445, 446, 452	108, 109, 112, 160, 207, 245
Jones, G. G., 26	Lapworth, A., xi, 26, 39, 104,
Jones, L. W., 370	149, 160, 151, 170, 180, 206,
Jones, W. H., 164	
Jones, W. J., 476	224, 225, 228, 232, 233, 238,
Jones, W. O., 446	242, 253, 267, 268, 273, 305,
Juliusberger, F., 341	317, 331, 380 383, 385, 386,
17 D	421,427,437,438,444,447,449
Karrer, P., 382	Larsson, E., 199
Kawamura, T., 193	Latiers, G., 138
Kekulé, A., xi, 3, 6, 12, 39, 106,	Latimer, W. M., 36, 212
117, 147, 405-8, 411, 412,	Laurent, A., 2, 195, 463
417-19, 422, 423	Lautsch, W., 119
Kellas, H., 283	Lavoisier, A. L., 1, 115, 195
Kendall, F. E., 111	Lawrence, C. D., 373
Kendall, J., 266	Lawrie, J. W., 110
Kenner, J., 224, 341, 359	Le Bel, 3, 290

Le Blanc, M., 85 Leermakers, J. A., 145 Le Fèvre, R. J. W., 342 Legard, A. R., 285 Lengfield, 364 Lenher, S., 132 Lennard-Jones, J. E., 15, 17-19, 47, 61, 66, 245 Lesslie, M. S., 342 Levene, P. A., 111, 309, 338 Lewis, G. N., x, 2, 12, 17, 29, 37, 100, 108, 120, 160, 168, 196, 200, 201, 225, 241, 419, 435 Liebig, J. von, 116, 288, 292 Lind, S. C., 85, 89 Lindemann, H., 108, 111, 112 Linstead, R. P., 179, 375 Lipscomb, A. G., 338 Lloyd, L. L., 284 Lobry de Bruyn, 265, 455 London, F., 15 Lonsdale, K., 406 Lorenz, L., 57, 400 Lorentz, H. A., 57, 400 Lossen, W., 7, 370 Löwenbein, A., 39 Lowry, T. M., xii, 23, 25, 26, 27, 37, 85, 109, 161, 197, 200, 205, 212, 265, 267, 269-72, 302, 304, 306-8, 311, 312, 316, 317, 326, 334, 336, 382, 387, 394, 399, 419-21 Lowson, W., 277 Lucas, H. J., 241, 390 Luchs, A., 327 Ludlam, E. B., 85 Lumsden, J. S., 179 Luther, R., 154 Luxmore, C. M., 361, 362 Maass, E., 171 Maass, O., 266 McBain, J. W., 274 Macbeth, A. K., 168, 224 McCombie, H., 265, 284, 286, 287, 478

McIntosh, D., 266

Marcus, E., 192

McNab, M. C., 179 Magson, E. H., 311

McKenzie, A., 327, 328, 338, 348

Mark, H., 72, 124, 157 [448 Markownikow, xiii, 178, 179, 376, Marsden, R. J. B., 395 Marsh, J. K., 422 Marshall, A. L., 132 Marshall, P. G., 130 Martin, A. R., 22 Martius, 475 Marvell, C. S., 124 Mason, J., 94 Masson, I., 450, 428 Mauguin, 364, 366 Maxwell, J. C., 56, 62 Mayo, F. R., 172, 179 Mecke, R., 14, 37, 44, 45 Meerwein, H., 138, 171, 231, 326, 350, 351, 359 Meisenheimer, J., 80, 352, 353, 454 Mendeléeff, 9 Merz, V., 479 Meyer, J., 280 Meyer, K. H., 124, 187, 189, 301, 303, 318 Meyer, L., 69, 79 Meyer, Lothar, 9 Meyer, Victor, 204, 206, 242, 283, 405, 471 Michael, A., 160, 179, 284,303,475 Mileska, L. A., 111, 338 Millet, H., 478 Mills, W. H., 30, 164, 340, 342, 359-61, 414, 417 Milner, S. R., 250 Moelwyn-Hughes, E. A., 83, 84, 246, 252, 271 Moest, M., 138 Montagne, P. J., 365, 366 Moore, T. S., 36 Moseley, H. G. J., ix Mosotti, O. F., 55, 59 Mosses, A. N., 249 Moureu, C., 90, 150 Moyer, W. W., 367 Mühlschlegel, H., 309 Müller, F., 112 Mulliken, R. S., 14, 18 Nathan, W. S., 246, 248, 254, 285 Nef, J. U., 107, 109, 110, 117,

142, 320 Newling, W. B. S., 271

Purkayastha, R. M., 144 Nernst, W., 86, 196 Pyman, F. L., 469 Nicholson, 1 Nicolet, G., 140 Race, E., 428 Nishida, D., 123, 187 Nixon, I. G., 417 Rahn, F., 434 Ramsperger, H. C., 145 Noddack, W., 149 Noelting, E., 425, 446, 479 Rancke, A., 149 Norris, J. F., 229, 230 Randall, M., 196, 225 Norrish, R. G. W., 26, 45, 144, Raoult, 61 Read, J., 174 148, 172, 173 Reed, J. B., 28, 29 Noyes, W. A., xiv, 111, 167 Reicher, L. T., 282 Reid, E. E., 263, 269, 284 Obermiller, J., 430, 433 Ochs, R., 193 Reiff, O. M., 151 Rice, F. O., 142, 143, 145, 273 Offenbacher, 124 Ogg, R. A., 176, 341 Richards, E. M., 312 Rideal, E. K., 151 Olivier, S. J. C., 227, 229-32, 235, 236, 237, 263, 285 Onoda, T., 132 Roberts, D. C. V., 362 Robertson, J. M., 406 Onsager, L.; 251 Robinson, P. H., 468 Robinson, R., xi, 39, 169, 170, Orton, K. J. P., 475-7, 481 Ostwald, W., 214, 229, 273, 285 171, 180, 182, 220, 240, 242, 252, 348, 369, 386, 388–91, Owen, G., 312 Oxford, A. E., 390, 392, 438, 444 410, 411, 415, 416, 418, 421, 438, 443, 444, 448, 451, 452, Palomaa, M. H., 281 454, 457, 472 Rodebush, W. H., 37, 212, 396 Paneth, F., 119 Pasiut, L., 176 Rollefson, G. K., 132 Rosen, R., 193 Pasteur, L., 290 Paul, T., 126 Rosenberg, I., 244 Pauli, W., 9, 18, 43 Roth, K., 125 Roth, P., 313 Pauling, L., 17, 24, 41, 112-14, Roth, W. A., 112 341, 394, 395, 407, 422 Rothstein, E., 243, 447, 449 Pechmann, H. von, 266, 322 Penney, W. G., 422 Perkin, Sir W. H., 210 Ruhemann, S., 408 Rutherford, E. (Lord), ix, 8 Perkin, W. H. (Jun.), 416 Rydberg, 9 Rydon, H. N., 179 Perrin, F., 89, 154 Pfeiffer, P., 36, 331 Phillips, H., xii, 31, 329, 330, 334, Sakellarios, E., 434 Samuel, R., 18, 35, 42, 48 338 Pickard, R. H., 328 Sandmeyer, T., 147 Scarborough, H. A., 265, 284, Pickles, N. J. T., 246 286, 287, 467 Piggott, H. A., 299 Plotnikow, J., 154 Scheele, C., 288 Polanyi, M., 86, 133, 269, 341 Schimmelschmidt, K., 464 Schlenk, W., 118, 123-7, 135, Pontin, 116 Porter, C. W., 324, 354, 475 157, 165, 192, 193 Schlenk, W. (Jun.), 165 Pounder, F. E., 428 Schlubach, H. H., 122 Pschorr, R., 147

Schmidt, J., 303

Purdie, T., 309

Schönberg, A., 321 Schroeter, G., 364, 370 Schumann, 397, 400 Schwarzenbach, G., 245 Scott, J. P., 309 Scarle, N. E., 342 Selby, W. M., 192 Sellmeier, 57 Seltz, H., 478 Semenoff, N., 91 Shapiro, U. G., 225 Shaw, F. R., 444, 447 Sherman, J., 394 Shoesmith, J. B., 228, 232, 233 Shoppee, C. W., 348, 452 Shriner, R. L., 339 Sidgwick, N. V., 18, 31, 34, 37, 45, 46, 65, 72, 108, 112, 113, 189, 255, 344, 396, 452 Simon, 307 Skrabal, A., 281 Skraup, Z. H., 413 Slater, R. H., 228 Smale, C. A., 94 Smallwood, H. M., 245 Smith, G. F., 269, 274, 444, 445 Smith, J. C., 390, 392, 438, 444 Smith, J. L. B., 414 Smith, J. M., 447 Smyth, C. P., 52, 58, 63, 64, 202, 203, 257, 258 Snethlage, H. C. H., 274 Snow, C. P., 130 Soper, F. G., 444, 445, 477 Southgate, H. W., 399 Spence, R., 143 Spencer, F., 446 Spong, A. H., 249 Sprankling, C. H. G., 284 Stahly, E. E., 326 Stanley, W. M., 343 Staudinger, H., 110, 146, 152-4 Steacie, E. W. R., 144 Steele, V., 306 Steiger, A. L. von, 257 Steiner, H., 47 Stelzner, R., 310 Stephens, H. N., 151 Stern, O., 60, 120 Stewart, T. D., 26, 172, 173

Stieglitz, J., 160, 354, 363-6, 370, Sudborough, J. J., 284 Sugden, S., xii, 28-30, 31, 37, 41, 42, 63, 108, 120, 126, 128 Sullivan, F. W., 124, 187, 191 Sutton, L. E., 58, 79, 112, 113, 255, 394, 395, 451 Tanret, 308 Tasman, A., 287 Taylor, H. S., 132, 164 Taylor, T. W. J., 362 Terry, E. M., 175 Thiele, H., 112 Thiele, J., xi, 107, 110, 145, 179, 372-4, 377-85, 389, 393, 412, 420, 462, 471 Thomas, W., 112, 113 Thompson, H. W., 91 Thompson, R. N., 94 Thomson, Sir J. J., xi, xiii, 10, 39, 54, 56, 57, 69, 76, 77, 85, 95-7, 381, 412, 418, 422 Thorpe, J. F., 216, 217, 310, 373 Tiemann, 364 Tietz, E. L., 150 Tietze, E., 475 Tiffeneau, M., 347, 348 Todo, 354 Topley, B., 341 Traill, R. C., 83 Tschitschibabin, A. E., 455 Tudhope, T. M. A., 338 Turner, E. E., 166, 342, 466, 468 Twitchell, 263 Udby, O., 272 Uhlenbeck, G. E., 16 Ullmann, F., 135, 458 Urech, 312 Usherwood, 317 (see Ingold, E.H.) Vakil, K. H., 480 Valasek, J., 59 Van der Kam, E. J., 464 Van der Waals, 55 Van Dorp, W. A., 366 Van Duin, A., 235

Van Itallie, T. B., 382

Van't Hoff, J. H., x, 3, 6, 8, 12, 30, 44, 100, 102, 290, 334
Van Vleck, J. H., 46, 120
Vargha, L., 321
Vass, C. C. N., 432
Volmer, M., 85
Vorländer, D., 205, 209, 426, 427, 436, 437, 448

Wachholtz, F., 139 Wagner, C., 132 Wagner, G., 350, 359 Wagner-Jauregg, T., 327, 345 Walchner, 288 Walden, P., 67, 118, 119, 124, 168, 187, 189, 195, 324-8, 330, 334, 338, 340, 341, 359 Walker, J., 137, 179, 289 Wallach, O., 365 Wallis, E. S., 340, 367 Warburg, E., 86, 327 Ward, D., 447, 449 Warren, F. L., 377 Waser, E., 401, 409 Waters, W. A., 85, 224, 245, 248, 253, 254, 265, 271, 451, 467-9 Watson, H. B., 149, 248, 254, 285 Wegscheider, R., 213, 280 Weickel, T., 126 Weigert, F., 154 Weissberger, A., 112 Weith, W., 479

Werner, A., xi, 3-7, 11, 22, 30, 34, 80, 100, 139, 331, 333,

341, 381, 429 Werner, E. A. 290

Wheeler, T. S. 94

Wheland, G. W. 394 Whiteley, M. A. 299 Whitmore, F. C., 94, 326, 349 350, 365-7 Widmer, R., 382 Wiede, 51 Wieland, H., 118, 124, 125, 145, 147, 153, 189, 230, 231, 355, 434 Wierl, R., 71 Wijs, J. J. A., 277 Wilkins, H., 28, 29 Willey, E. J. B., 89 Williams, E. G., 246, 247 Williams, Gwyn, 89, 173, 477 Williams, J. W., 64, 396 Williamson, A., 2, 135 Willstätter, R., 373, 401, 408, 409, 416 Wilson, I. S., 260 Winmill, T. F., 36 Winterstein, A., 382, 397 Wislicenus, J., 106, 138 Wohl, 454 Wöhler, 116, 288, 289, 292 Wolf, K. L., 58, 139 Wolfrom, M. L., 309 Wollschnitt, H., 192 Wolter, A., 111, 112 Wrede, E., 61 Wright, C. H., 171 Wurtz, C. A., 135 Wyckoff, R. W. G., 42 Wynne-Jones, W. F. K., 197, 199, 200

Young, J. H., 339 Yuan, H. C., 342, 345, 346

Zahn, C. T., 79, 395 Ziegler, K., 124, 157, 192 Zincke, T., 297, 483 Zintl, E., 127 Zoellner, E. A., 192

SUBJECT INDEX

Absorption spectrum, 9, 130, 190, 397-9, 421, 422 Acetoacetic ester, acidity of, 204, 205; tautomerism of, 300-302, 305; alkylation of, 391 Acetone, 25, 26, 177; substitution of, 305, 317, 318 Acetyl, free radical, 146 Acetylene, 27, 49; derivatives, Acidity, 195-220; induction of, 204-20; nature of, 195-9; origin of, 201-3 Acids, as catalysts, 264, 267-79, 312-16, extended definition of, 196-9; ionic reactivity of, 165-6 Activation, 26, 81-3, 86-8, 161, 358, 383–8, 403, 419–21 Active nitrogen, 88, 89, 120 Activity coefficient, 196, 277 Acrylic acid, addition to, 179, 180 Acylous group, 205-13, 224, 232, 233, 236, 239-43 Addition reactions, ionic, 170-**.** 180, 240, 266-9, 413, 433-5, 439; non-ionic, 102-6, 132, 141 Aldehydes, 150-2, 169 Aldoximes, dehydration of, 352, 353; stereoisomeric change of, 361, 362 Aliphatic diazo compounds, 110-114, 153 Alizarin, formation of, 455-6 Alkylating agents, 166, 167, 181, Alkyl halides, 162, 166, 167, 221-238 - isocyanates, 320

Alkyl radicals, migration of, 320-71 --- sulphates, 166-7 Alternate polarity, 228, 380, 383-388, 477 Alternating affinity, 429–31 Amidines, 322, 323 Amino-crotonic ester, 170, 390 Ammonium radical, 115, 116, 122 Aniline salts, substitution of, 260, 446, 447, 449, 450 Anhydronium bases, 415, 416 Anionoid reagents, 180-5, 439, 447, 450, 454, 455 Anionotropic change, 318-20, 322, 323, 387, 474 Anode reactions, 137, 138 Anthracene, 130, 131, 141, 154, 414, 420, 434, 435 Anti-knocks, 90, 152 Antoxygens, 151, 172 Aromatic sextet, 410, 411 — structure, 405–24 substitution, 228, 260, 425–83 Aryl radicals, migration of, 320-4 Aseptic distillation, 301, 318 Association, 64-6, 95 Atom polarization, P_A , 58, 60 Atomic structure, ix, 8–10 — orbital, 17–19, 48 — reactions, 84, 85, 115-58 refraction, 400 Auto-catalysis, 264, 277 Autoxidation, 90, 150-2 Azides, 110–14, 363–9, 396 Azo benzil, 364, 365, 370 decomposition compounds, of, 144, 145 - methane, 145, 370

Baever's strain theory, 102; application to benzene, 417 Band spectra, 12, 14 Cf. Basylous group, 205-13. Acylous group Beckmann transformation, 80, 325, 326, 351-61; analogues of, 363-7 Benzidine transformation, 481 Benzil, 349, 373, 374, 393; oximes of, 353, 362 Benzoyl chlorides (substituted), hydrolysis of, 230 Benzyl halides (substituted), hydrolysis of, 227-9; reduction of, 232 Betaines, xii, xiii, 52, 62, 112, 209 Bivalent carbon, 42-6, 106-10, 117, 132 - hydrogen, 35-7 — nitrogen, 124-6 Bohr theory of line spectra, 13, 14 - magneton, 120 Boltzmann's constant, 54, 250 Bond polarizability, 253, 256-9 - polarization, 253, 254 - refractivity, 257, 258 Boron, fluoride, 6, 32, 33, 200; hydrides, 5, 40, 41 Branched chain reactions, 90-2 Bromination, 172-6, 305, 372, 373, 435, 442, 443, 464, 465 Bromohydrins, 174-6 Butadiene, 372, 373; derivatives of, 386, 389 Butylene, iso-butylene, 290-2 Camphor, 300, 398, 399; bromo-, 306; nitro-, 223, 224, 306 Camphene, 171, 351 Carbon, ³P and ⁵S forms of, 43, 46; amphoteric character of, 185–93 Carbonium ions, asymmetric, 339, 340; rearrangements of, 326; stability of, 355 - salts, 156, 186-92, 461 Carbon monoxide, reactions of, 86, 88, 106-9, 132, 148; structure of, 43-6

Carbonyl group, activation of, 178; dipole moment of, 178; kationoid reactivity of, 180, 184; photochemistry of, Carboxyl group, polar effects of, 214-8, 240; structure of, 36, 205, 392, 395, 396 Catalysis, by electrical fields, 358, 362; of hydrolysis, 159, 264-77; of isomeric change, 295, 296, 305, 312–16; polar, 82, 94, 98, 159; by proton, 198, 199 Catalytic catenary, 276, 277 Centric benzene formula, 408-11, 418 Chain reactions, 132, 133, 139, 151-5, 173 Chelate rings, 411 Chemi-luminescence, 133-5 N-Chloramines, 475–8 Chlorobenzene, substitution of, 443, 452 Chlorohydrins, 174–6 Cinnamic acid, addition to, 132, 139, 140, 176; substitution of, 426, 428, 437 of conjugated Classification, systems, 388-93; of reactions, 243, 244; of reagents, 180-5 Clausius-Mosotti equation, 54, 59 Collision frequency, 82-4, 134, 248, 250, 357, 358 Colour of aromatic kations, 190-2 Compressibility of radicals, 345 Conjugation, xi, xiv, 170, 219, 220, 259, 316, 37 - 404, 411-418, 423 Co-ordination, 5, 30-42, 163, 231, 237, 379 Coupled electrons, 12, 15, 394 Covalence, xi, 11, 12, 17, 21, 25, Crossed polar systems, 384, 393, 470 Crotenoid systems, 390 Crotonoid systems, 392 Crum Brown-Gibson rule, 426,

427

Cumulative double bonds, 401 Curtius's reaction, 147, 363-70 Curved arrow symbollism, 386 Cyclo-Octatetrene, 416. 402, Cyanhydrin reaction, 26, 34, 104, 105, 170, 236, 267, 268 Cyanogen, 115, 116 Cyclo-Hexadiene, 397, 398, 402, 403

Dative bonds, 31-4, 38, 163 Debye's polarization equation, 52-5, 63 Deformation of ions and molecules, 21, 22, 51, 68, 251, 344, 345 Degree of dissociation, 22, 187–8 Desmotropy, 300, 303, 304 Detonation, 152 Dewar's benzene formula, 408 Diamagnetism, 47, 120 Diazenes, 146, 147 Diazo-acetic ester, 146, 147, 369, Diazo-amino compounds, rearrangements of, 475, 476 Diazo compounds, 147, 181, 204, 205, 319 Diazomethane, 144, 369 Dibasic acids, dissociation constants of, 213-18, 229, 247, 248; ester hydrolysis of, 281, 282 iso-Dibutylene, iso dibutol, 293-~ 295, 317 Dielectric constant, 52-64, 97, 98, 200, 251, 252 Diels-Alder reaction, 376, 377, 414, 472, 473 Dimensions of radicals, 343, 345 Dinitro-chlorobenzene, 458; -toluene, 459, 460 Diphenyl (and derivatives), II. 70; optically active, 341-6; substitution of, 260, 465-9 . Dipole moments, xiii, 21, 37, 38, 50-80, 108, 109, 161, 208, 341, 343

Dipoles, chemical influences of,

239-50, 341, 347-50, 358, 359, 363, 448, 451 Direct effect, D, 253 Directed valency, 6, 12, 102, 333 Directive power (in aromatic substitution), 431, 432, 467, **468** Discordant polar systems,—see Crossed polar systems Displacement reactions, atomic, 136-8; ionic, 226, 227, 236, 237, 481 Dissociation constants (tables of), 207-13, 218-20, 229, 247, 248 Distributed charge, 393-6 Donors and acceptors, 31-4, 108, 181, 184 Double bonds, 25-30, 99-106 Dualistic theory (of Berzelius), x, 1, 3, 7, 8 Dual theory of catalysis, 274 Dyad systems, 316, 317, 319 Dynamic isomerism, 290-7, 302, 304 Einstein's law of photochemistry,

85, 140, 143, 149 Electrical conductivity, 22, 93, 124, 127; measurements of, 321, 322 Electric circuit (in reactions), 169, 194, 443 Electrode potentials, 225; reactions, 137, 138 Electrolysis, 1, 4, 137, 138, 159 Electromeric change, 170, 388, 391; in benzene ring, 411-461, 482, 483; in naphthalene ring, 464, 465 (also = electrotropy, electromerism) Electron, affinity, 182-3; displacement, 51, 53; drift, 241-4, 246; diffraction, 110;

> duplet, 14-17 (= covalent emitting,

> group and sub-group, 9, 10,

12; mass and charge of, ix, 8; migration, see electro-

merism; orbit, 12, 18-24,

203; planetary, 8-10; sex-

14-17;

bond);

```
Electron (continued)
    tet, 32, 33; sharing, x, xi, 8-
    12, 31-4; source and sink,
    242, 278, 392; spin, 121;
    transfer, 8-11
Electronic energy, 13-17; ex-
    citation, 87; polarization = P<sub>E</sub>, 58, 60; structure of
    atoms, 8, 9; of benzene, 39,
    418-23; of molecules, 17, 18;
    theory of valency, 8, 10-12
Electrostatic
              fields,
                        II.
    chemical influences of, 96,
    119, 234, 239-54, 440-1,
    447-51; symbolism of, 243,
    249
 - induction, 241, 249, 252-4
Electrostriction, 22, 216, 241,
    250-2
Electrovalence, xi, H, 12, 18-21,
    25, 26
Energy, of activation, 82-4, 225,
    245-7, 285; of bond for-
    mation, 17; chains, 89; of
    ionization, 19-21; levels,
    14, 16
Enols, 27, 301, 302, 314, 325,
    377-9, 391
Erlenmeyer's rule, 378
Esterification, 81, 159, 272, 330,
Esters, hydrolysis of, 262-87;
    structure of, 396
Ethyl, radical, 118
Ethylene, reactions of, 26, 129,
    132, 171-7; structure of,
    25, 29, 130
Exchange, energy, 12, 40; forces,
    121
Excited forms, of atoms and
    molecules, 13, 43, 45, 47, 88,
    89
Explosion waves, 85, 86, 90 •
Faraday tubes of force, 420, 429,
Fine-line structures of spectra,
    14, 87
Fischer-Speier reaction, 283
Fission test for tautomerism, 299
Fluorescence, 88, 89, 155
```

Fluoride, boron, 6, 32-4, 200; potassium, structure of, 18, 35
Flutings of band spectra, 13
Free affinity, 156, 157, 407, 412
— proton, 197, 198
— radicals, xiv, 39, 40, 99, 115-58, 462
— rotation, 54, 71-3, 78-9, 100, 139-41, 341-6, 367
Friedel-Crafts reaction, 230-2, 235, 434, 476
Fry's benzene theory, 419, 435, 436
Fulminates, 106, 117
Functional tautomerism, 299

Gas reactions, 82, 85-92
General polarity, 238-61, 263, 440; definition of, 238-9; effects of, 246-61, 278-87; nature of, 239-46
Glucose, mutarotation of, 307-16
Glycerates, hydrolysis of, 279, 280

Graphitic conduction, 387 Grignard reagents, 164, 165, 186, 225

Halogen compounds, polarizability of, 257-8; reactivity of, 191, 207, 208, 219, 221-37
Halogens, 10; ionic reactions of, 92, 149, 167-84; non-ionic reactions of, 138-41, 148-50; polarizability of, 257, 258

Heat of bond formation, 44-6, 102; of combustion of hydrocarbons, 49, 102, 408; of conjugation, 377, 393, 395 Hell-Volhard reaction, 149 Hetero-enoid systems, 390, 391, 442-6, 461, 474 Hexaphenylethane, 117; see Triphenylmethyl

Hofmann degradation of amides, 363-8; rearrangement of amines, 481

Hollemann's substitution rules, 431-2

Hydrindene, 417, 418 Hydrocarbons, absorption spectra of, 397; molecular refraction of, 400 Hydrogenation, ionic mechanism of, 176, 177 bond, 18, Hydrogen 34-9; chloride, 19, 20; cyanide, 27, 44, 296, 317, 318; ion activity, 195, 196; molecule, 15-17; peroxide, 27, 138; see also Acids and Proton Hydrolysis, 227-30, 234-7, 243, 244, 262-87; catalysis of, 264, 265, 273; mechanism of, 263-72; rate of, 270, 272-82, 285-7 Hydroxylamines and Hydroxamic acids, rearrangements of, 370, 371 Hydroxyl radical, 40 Hypochlorous acid, 168 Ice, structure of, 35–8 Imino-aryl ethers, 321-23 Imported polar effect, 253 Indirect substitution, 328-31; (aromatic), 428, 429 Induced dipole moments, 76, 253, 255, 256, 259 - polarity, theory of Lapworth, xi, 437, 438; see also General polarity Induction of ionic reactions, 95-7 - period, 151, 173 Inductive influences, of substituents, 204-61, 394, 441; Ingold's symbolism for, 242, 243 Inhibition of photochemical reaction, 155 Interatomic distance, 15, 215-18, Interconversion of geometrical isomers, 138-40, 361-3 Intermolecular rearrangement, 476-80 Intramolecular ionic charges, 22-24, 27 - rearrangement, 471, 481-3; compare chap, xiv

Intrinsic acid strength, 200 Inverse substitution, 222-5, 325 Iodine monochloride, 168 Iodo dichlorides, rearrangement of, 478 Ionic, dissociation, 22, 93; linkage, see Electrovalence; reactions, 84, 92-8, 104, 105, 159-94; structure, 10 Ionization, potential, 182; reagents, 162-8 Ionotropy, 288–323, 414–16, 474– 483 Ion-pair, 21, 22, 50, 52, 83, 93, 277 Inverse substitution, 168, 222-5, Iso-Cyanates, 289, 363, 364 *Iso-*Cyanides, 106–9, 111 Isomeric change, 171, 264, 269, 288–323, 356–60, 389, 428, 429, 473, 474 Isoterism, 27, 43, 107, 148 Isotypic series, 247, 278 Katio-enoid systems, 392, 453-59, 461, 470-3 Kationoid reagents, 171, 180-5; reactions (aromatic), 439, 440, 442, 450, 453-9 Ketens, 152, 153 Kekulé's benzene formula, xi, 405-8, 417-19, 422, 423 Key atom, 438 Kolbe's synthesis, 137, 138 Lapworth's $\alpha \beta$ rule, 317; $\alpha \gamma$ rule, 380, 383, 385, 389, 438 Lead ethyl, 90, 118, 152, 164 Lone pair of electrons, 31, 67, 200, 201 Lorentz-Lorenz equation, 57 Lossen rearrangement, 370, 371 Lowry's benzene formula, 419-21 Magnesium hydride, MgH, structure of, 14, 15 Magnetic coupling between electrons, 17; moment, 54, 120,

150; properties of unpaired

electrons, 40, 47, 119-21;

susceptibility, 120, 121

Maleic anhydride, 376, 377, 393, β Naphthol, bromination of, 435, 464-5 470 Malic acid, 309, 326, 327 Malonic acids (substituted), 217, Markownikow's rule, xiii, 178-80, 376, 448 Maxwell's law of probability, 62, 98 Mercuric salts, 20, 32, 34; alkyls, 164, 192 Merotropy, 303 Mesomeric state, 305, 314, Metallic alkyls, 163-5; ketyls, 121-9, 136, 157 Meta substitution, mechanism of, 449, 450 Methyl, anion, 347; azide, 113, 305; free radical, 117, 119, 204 133, 134, 138, 143, 145; glucosides, 300, 308; pyridines, 460 Methylene radical, 144 Michael's distribution principle, 147 Migration of alkyl groups, 320, 476, 481-3; of double bonds, 378 Mixed double bond, xii, xiii, 26-28, 30; compare Semi-polar bond Molecular beams, 51, 60, 61; electronic orbitals, 17-19; orientation, 250; polarizability, 52-5, 403; polarization, 52-63, 256; arrangement, 94, 324-71, 473-83; refraction, 28, 57, 202-6, 256-8, 399-404; rotation, 87; vibration, 87; volume, 28 Moloxides, 152, 153 Molozonides, 153, 154 Multiplicity of line spectra, 15 Mutarotation, 306-15 Naphthalene, formula of, 11, 409, 413, 420; substitution of.

462-5; sulphonic acids, 479-

48I

Negative catalysis, 89, 90, 151, Neutral salt action, 273 Neutralized systems, 392, 415, 471, 472 Nitramines, rearrangement of, 475, 481 Nitration of bases, 252, 446, 447; of camphor, 306; of olefines, 433, 434 Nitric acid, structure of, 165, 166 — oxide, 12, 40, 48, 120-3, 125, 138, 150, 155 Nitro-acids, 205, 206; -benzene, 453-5; 447, 449, 450, -butane, 339, 340; -camphor, 223, 224, 307, 311; -diphenyl, 466; -paraffins, -group, properties of, 208, 209, 467-9; structure of, 73, 74, 166, 396 Nitrogen peroxide, 12, 40, 122, Nitrosamines, rearrangement of, Nitroso-dimethylaniline, -phenol, 297-9, 414, 474 Nitrous oxide, 27, 66 Non-ionic reactions, 84-92, 105, 106, 115-58 β -Octanol, 337, 338 Octet rule, 29, 41 Odd electron bónds, 39-42, 100; -reagents, xiv, 48, 120-58 Olefines, ionic reactions of, 170-177, 433, 434; non-ionic reactions of, 102-6, 138-41 Optical exaltation, 401-4; inversion, 307-10, 324, 326-41; rotation, 329 Orbital coupling of electrons, 47 Organic anions, 163-5, 192, 193; kations, 162, 163, 186-92 Organo-metallic compounds, 181, Orientation laws for benzene, 425-8

Oxidation, 151, 154, 413 Oxidizing and reducing agents, 33, 181, 183, 184 Oximes, 30, 79, 80, 101; rearrangements of, 325, 351-Oxonium compounds, 165, 174, 266-8, 317; ions, 20, 21, 28, 35, 196-9, 201, 203 Oxygen, reactions of, 32, 89-92, 150-2; structure of, 46-9 Ozone, 27, 49 Parachor, xii, 28-30, 41, 108, 109, 112, 153 Paramagnetism, 33, 47, 48, 101, 120, 121 Partial valency, xi, 381, 382, 385, 412, 420 Pauli's exclusion principle, 9, 18, 20, 43 Pentaphenyl ethyl, 124, 157 Per-acids, 150, 151; -bromides, 42; -oxides, 27, 49, 90, 91, 126, 129, 146, 150-3, 172, 179, 393 Periodic system, 9, 16 Perturbed structures, 254, 305, 394-6 Phenols, acidity of, 204, 208; ethers of, 475; substitution of, 442-6, 483 Phenyl -acetic acid, 210; -allyl alcohol, 319-20; anion, 322, 349, 350; -butadiene, 389; group, 339, 461; -hydroxylamine, 482; -nitromethane, 437; -pyridines, 469; radical, 145 Phosphorus pentachloride, 41 Photochemical threshold, 87; reactions, 85-90, 123, 129-32, 138-44 Pinacones, 127, 146, 177, 374; rearrangements of, 325, 346-348 Pinene, 152, 171, 350 Polar-activation, 26, 104-6, 222, 382-8, 403; -catalyst, 26, 98, 173, 174; fields, effects of, 21, 93, 238-61, 278-87,

350; reagents, 26, 98, 159-94, 336-9 Polarizability, bond, 256-61; of compounds, 202-6, 213, 257-61; molecular, 52-5 Polarization, 20-2, 52-65; in benzene ring, 438, 441, 448; of halides, 234; molecular, 55-63, 256 Poles, field due to, 244, 245 Polycyclic hydrocarbons, 414 Poly-noid systems, 389, 414 Polyhalides, 41, 42 Polymerization, 119, 141, 142, 153, 376, 377 Positive halogens, 167, 168, 220-226, 232, 233, 302 Predissociation, 86, 87 Primary Valency, 4, 5, 34 Propylene, addition to, 179 Proton, 8, 19, 20, 32, 38, 39, 213, - donors and acceptors, 196-8 Prototropic change, 198, 199, 206, 300, 304, 306-16, 355, 356, 387, 410, 456, 457, 474 Pseudo-bases, 320; -halogens, 120-2; -metals, 122 Pseudomerism, 298, 317 Pvridine, as catalyst, 314, 315, 337, 339; reactions of, 455 Pyridones, 456, 474 Pyrrole, 410, 457 Quadrivalent nitrogen, 124-6 Quantum, of light, 87; mechanics, 13, 15, 82; number, 9; theories, 7-9, 19, 120, 121 Quinoles, 482 Quinoline, derivatives, 455-6, 460, 465 Ouinones, 393, 470-3 Quinonoid structures, 190-2, 454, 455, 474 " R " form, 131 Racemization, 67, 329, 331, 336, 341-€ Rate of aromatic substitution,

432, 433, 449

Reaction constant = \mathbf{B} , 82-4, 134; chains, 85-92; velocity, 82-4,228-32 Reduction, 177, 232, 233 Refractive dispersion, 399, 400 Reimer-Tiemann reaction, 483 Residual affinity, 372, 380-2, 413, 430; -valency, 381 Resonance, 37, 113, 114, 157, 305, 382, 393-5, 421-3, 462 - spectrum, 134 Resonating electrons, 40 Reversed field effect, 249 Ring-chain tautomerism, 310 Rotational energy of molecules, 3, 4, 6, 14 Rydberg series, 9, 10 s and ϕ electrons, 17, 18, 43 Sandmeyer's reaction, 147, 148 Saponification, 92, 264-87 Secondary alcohols, Walden inversion of, 328-30, 334, 335 - valency, 5 Semihydrobenzoin rearrangement, 348 Semipinacolinic rearrangement, 348 Side-chain migrations, 474-8, 481-3; polar effects of, 260, 261; reactivity of, 459, 460 Silicols, 101 Semi-polar bond, xii, xiii, 28-31, 41, 44, 52, 101, 103, 111, 208, 314 Single electron bond, 39-42, 128 Sodium metal, reactions of, 86, 105, 123, 129, 133-6 - alkyls, 163, 164, 192 — triphenylmethyl, 192, 193 Solvation, 22, 189, 252 Solvents, effects on reactivity, 154, 193, 194, 199, 200, 213-218, 265-72, 311-16 Spectra, atomic, 13-15; molecular, 14-15 Spinning electrons, 12, 16, 17, 40, 47, 120 Stability of conjugated systems, 372, 377, 379, 393 Stationary states, 14, 88

Steric hindrance, 230, 283, 284, 336, 341-6 Stieglitz rearrangement, 370-1 Strong electrolytes, 251, 282 Substitution, aromatic, 228, 260, 425-83 Sugars, mutarotation of, 307-16 Sulphonic acids, 479-81 Surface, catalysis, 98; reactions, 89, 91, 173-4; tension, 28, 31 Symmetry test for tautomerism, Tautomeric effect = T, 388, 441, 461, 462 Tautomerism, 113, 114, 224, 297-320, 322 Tercovalent carbon, 334, 335 Tertiary amines, rearrangement of, 475 Tetraphenyl-hydrazine, 156; -methane, 145 Thiele acetylation of quinones, 471, 472 Thiocarbonic esters, 321 Thionyl chloride, in Walden inversion, 327, 328, 337-9 Thomson's benzene formula, xii, 39, 412, 418, 422 Three-body, collisions, 88 Three election bonds, 39, 128, 381, 382, 412 Toluene, substitution of, 94, 447, 449, 458 Trans-Addition, 140; migration, 352-4, 360, 361 Triad systems, 316, 319 Trinitrotoluene, 458, 459 Triphenyl isoxazole, 351, 352 Triphenylmethyl, 118, 119, 122-5, 127, 130-2, 145, 146, 150, 155-7, 186-91, 396 - chloride, 156, 186-91, 210 - peroxide, 123, 146, 335 Types, theory of, 2, 3

Ullmann's reaction, 136, 458 Univalent nitrogen, 364-8; oxygen, 126, 127 Unitary theory of affinity, 2, 7, 8 Unpaired electrons, 48, 49, 99, 368
Unsaturated acids, 211, 212, 218-220, 375, 376
— compounds, reactions of, 99-115, 184, 185
— ketones, 374, 375, 399
— hydrocarbons, 397, 398, 400
Urea, 288-90

Valency, deflection, 77-9, 216, 217, 344, 346; electrons, 12, 17; nature of, 5-12, 25-49 Van der Waals' equation, 55; forces, 61 Vector addition of dipoles, 69, 73-5, 77, 395 Vibration energy, 13; frequency, 44, 45, 47, 48 Victor Meyer's esterification law, 283, 284 Virtual Tautomerism, 299

Vorländer's substitution rule, 437

Wagner-Meerwein rearrangement, 350, 351

Walden inversion, 67, 324, 326-41; mechanism of, 331-6

Water, as catalyst, 173, 174, 312-14; structure of, 20, 37-9

Wave mechanics, 12, 21, 37, 113, 393

Werner's theory of Valency, 6, 7, 429

Williamson's ether synthesis, 83, 135, 136

Wurtz reaction, 136, 137

X-ray spectra, 9; diffraction, 71; structure of benzene, 406, 407

Zinc methyl, 163, 164

DATE OF ISSUE

This book must be returned within 3/7/14 days of its issue. A fine of ONE ANNA per day will be charged if the book is overdue.

will be char	ged if the b	ook is over	due.	
			1	Ī
l		1	ł	1
		1		- 1
		1	1	- 1
		1	1	1
				1
i		100		
1				į
-		1		1
	1			Ì
1	1			l
1				į
				- 1
4		1	1	- 1
1	361	1	1	
		1	1	- 1
			1	
1		1		

SEVEN DAY BOOK